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SCIENCE 10

Module 1

Energy and Matter in Chemical Change

Learning Technologies Branch

Alberta EDUCATION

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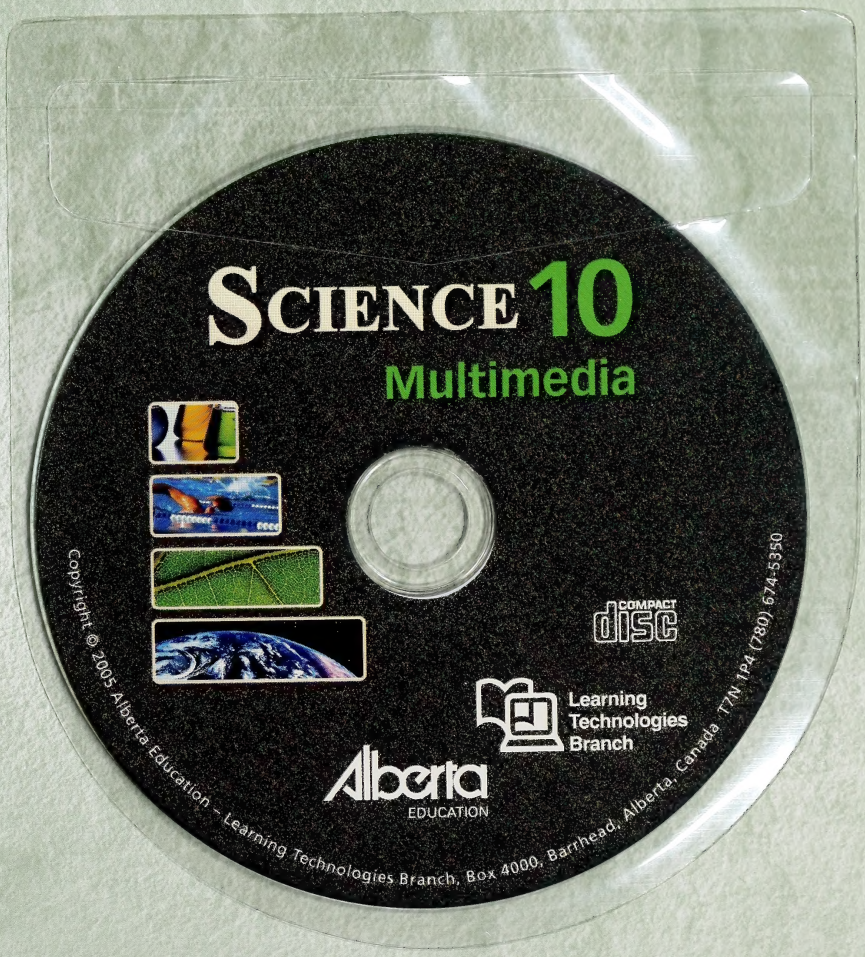
Multimedia

CD

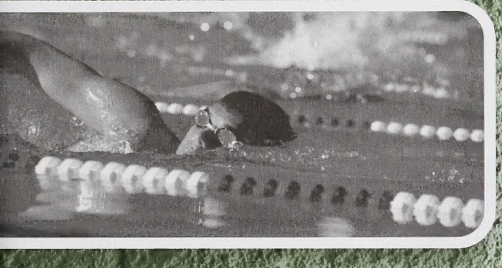
SCIENCE 10

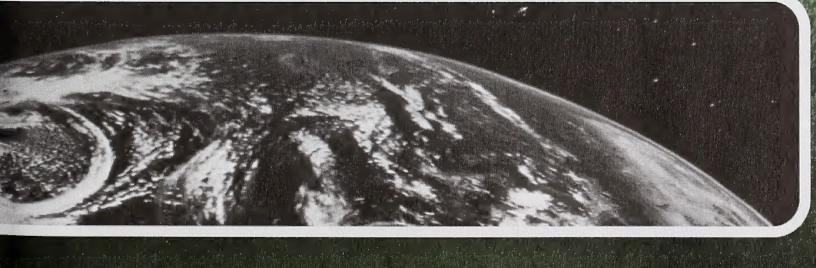


SCIENCE 10

Module 1

Energy and Matter in Chemical Change





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Science 10
Module 1: Energy and Matter in Chemical Change
Student Module Booklet
Learning Technologies Branch
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The Learning Technologies Branch acknowledges with appreciation the Alberta Distance Learning Centre and Pembina Hills Regional Division No. 7 for their review of this Student Module Booklet.

This document is intended for	
Students	✓
Teachers	✓
Administrators	
Home Instructors	
General Public	
Other	

You may find the following Internet sites useful:

- Alberta Education, **http://www.education.gov.ab.ca**
- Learning Technologies Branch, **http://www.education.gov.ab.ca/ltb**
- Learning Resources Centre, **http://www.lrc.education.gov.ab.ca**

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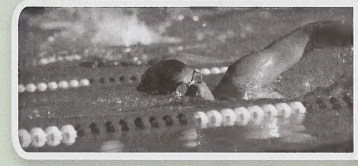
Welcome to SCIENCE 10

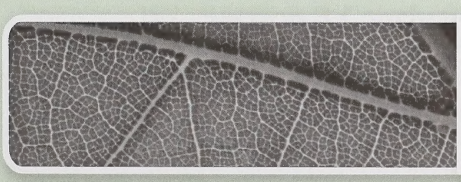
Module 1

It is recommended that you work through the modules in order (from 1 to 4) because concepts and skills introduced in one module will be reinforced, extended, and applied in later modules.

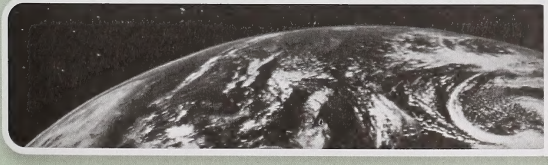
Module 1
Energy and Matter in Chemical Change



Module 2
Energy Flow in Technological Systems



Module 3
Cycling Matter in Living Systems



Module 4
Energy Flow in Global Systems

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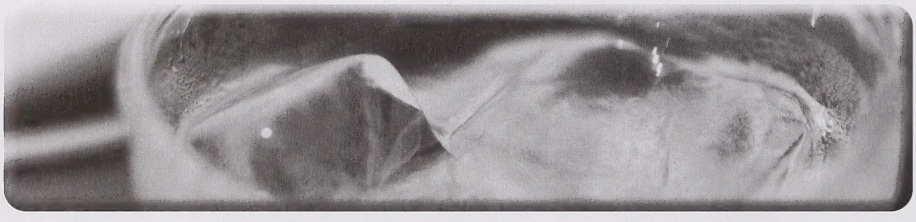

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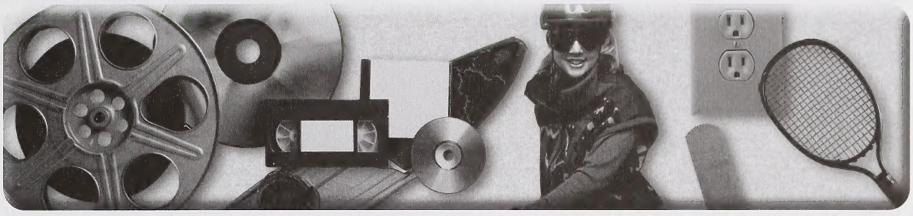

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T1 **Template for Inquiry Lab: Ionic or Molecular?**

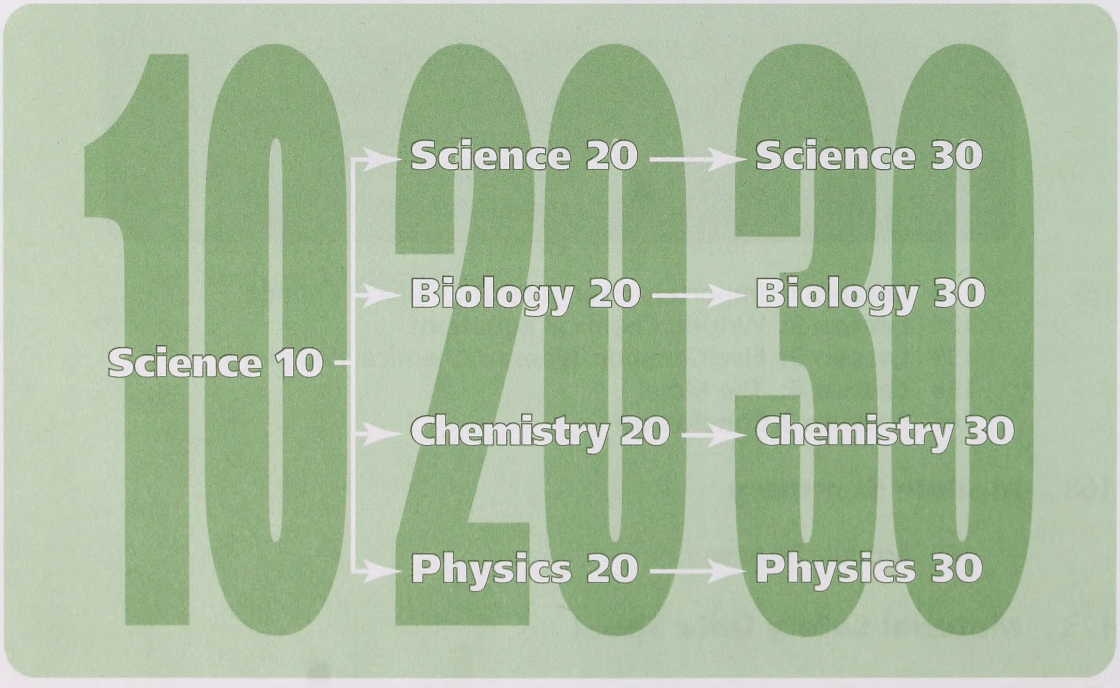
The Secondary Science Program

The Secondary Science Program is designed to develop your scientific literacy. Scientific literacy includes a thorough knowledge of science and its relationships to technologies and society. To become scientifically literate, you need to develop broad-based skills required to identify and analyze problems; to explore and test solutions; and to seek, interpret, and evaluate information. You also need to become aware of the role of science in responding to social and cultural change and in meeting needs for a sustainable environment, economy, and society.

The Secondary Science Program incorporates Aboriginal perspectives in order to develop an appreciation of the cultural diversity and achievements of First Nations, Métis, and Inuit peoples.

Learning opportunities are made meaningful so you can relate science to your life out of the classroom. This encourages an interest in science as a lifelong learning experience.

The secondary science course sequence allows for general science development through the Science 10–20–30 sequence or a subject-specific sequence through Science 10 and any of Biology 20–30, Chemistry 20–30, or Physics 20–30. The following graphic shows the course sequences leading from Science 10.



To support your development of scientific literacy, the Secondary Science Program provides a foundation of learning experiences to address the critical aspects of science and its applications. Four foundations provide the general direction for Science 10 and the courses that follow:

- **Science, Technology, and Society (STS).** You will develop an understanding of the nature of science and technology, the relationships between science and technology, and the social and environmental contexts of science and technology.

- **Knowledge.** You will extend your knowledge and understanding of concepts in life science, physical science, and Earth and space science.

- **Skills.** You will develop the skills required for scientific and technological inquiry, for solving problems, for communicating scientific ideas and results, for working with others, and for making informed decisions.

- **Attitudes.** You will be encouraged to develop attitudes that support the responsible acquisition and application of scientific and technological knowledge to the mutual benefit of self, society, and environment.

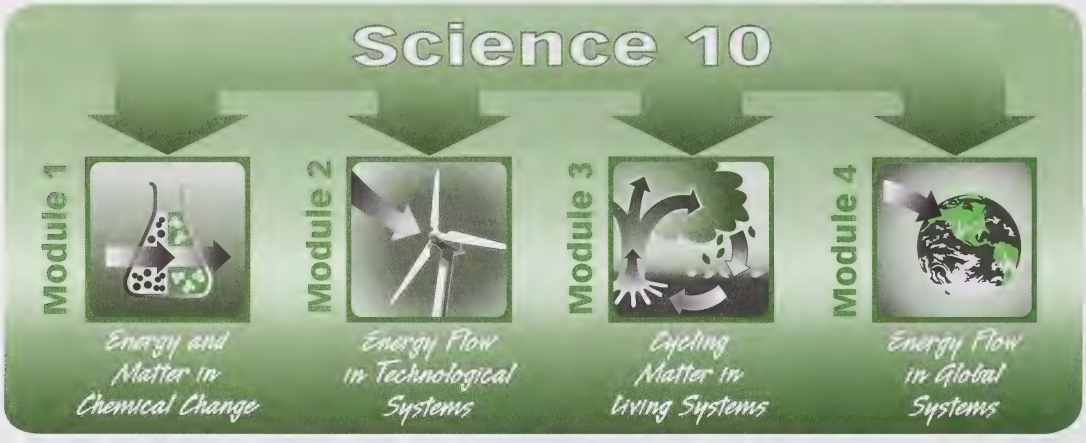
Consult your teacher or counsellor for information on Science 10 and any of the other secondary science courses in the sequence. You can find out more about the secondary science courses and high school requirements at the Alberta Education website.

http://www.education.gov.ab.ca

Overview of Science 10

Science 10 is the first course in the secondary science sequence. Students who successfully complete Science 10 may take Science 20, Biology 20, Chemistry 20, or Physics 20. In this course you will deal with four different areas of science: transformations of matter in chemical changes, energy flow in technological systems, cycling of matter in living systems, and energy flow in global systems.

This Science 10 course consists of four modules. Each module consists of one Student Module Booklet and three Assignment Booklets. The booklet you are presently reading is called a Student Module Booklet. It will show you, step by step, how to advance through each module.



Resources

Textbook

To complete the course, you will need the *Addison Wesley Science 10* textbook.

Multimedia

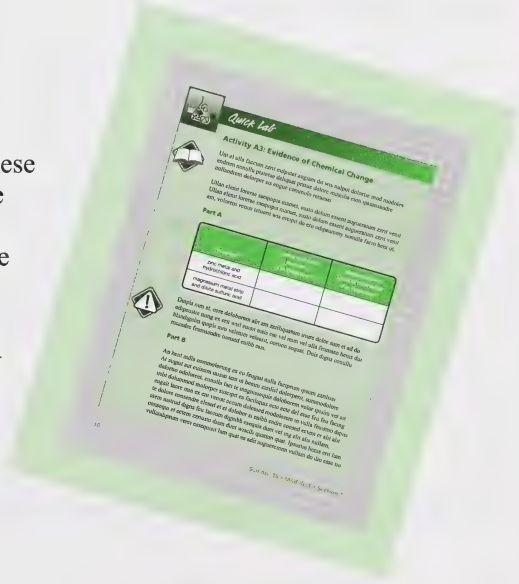
Attached to the inside cover of Module 1 of this course is the Science 10 Multimedia CD. This CD contains multimedia segments, photos, and spreadsheets designed to help you complete activities and better understand particular concepts presented in this course. Ask your teacher, friend, or family member if you need help using this CD.

Binder or Notebook

Because there are no response lines provided in the Student Module Booklets, you will need a loose-leaf binder or notebook to answer questions and complete charts. It's important to keep your lined paper handy as you work through the material and to keep your answers together in a binder for review purposes later.

Materials and Apparatus

The modules include numerous science labs and investigations that require materials and apparatus. These items are needed to complete the module. Some of the materials and apparatus may be provided at your local school lab. If you don't have access to a school lab, the investigations are set up with a Part A and a Part B. Part B will provide an alternate pathway for those without access to the materials and apparatus required for Part A.


Quick lab

Activity A3: Evidence of Chemical Change

Part A

Part B

Other Resources

The Internet

The Internet can be a valuable research and learning tool. Periodically, references to Internet sites are provided in the Student Module Booklets. **Note:** It is encouraged that you also use other information sources, such as your local library, to supplement your research.

When using the Internet, there are a few things to keep in mind:

- **Do not believe everything you read.** The Internet is filled with information. Unfortunately, not all of it is correct. Anyone can put information on the Internet. The important thing is that you take a close look at the source, that is, who is credited with supplying the information. For example, did a university, a museum, or a science centre put the information on the Internet? If so, it probably provides correct information.

- **Use online research tools.** There are a number of different tools, called search engines, available to assist you in finding information on the Internet. These engines organize and sort information by topic or key word. AltaVista Canada (**www.altavista.com**), Google Canada (**www.google.ca**), and Yahoo! Canada (**www.yahoo.ca**) are just a few of the search engines available.

- **Websites can change or disappear.** Keep in mind that the Internet is constantly changing. Sometimes you will discover that after typing in an address, a note will come up on the screen indicating that the site has moved or disappeared. If you find that an address does not work, go back to one of the search engines and do a general search using a key word or phrase.

LearnAlberta.ca

LearnAlberta.ca is a protected digital learning environment for Albertans. This Alberta Education portal, found at **http://www.learnalberta.ca**, is a place where you can support your learning by accessing resources for projects, homework, help, review, or study.

For example, LearnAlberta.ca contains a large Online Reference Centre that includes multimedia encyclopedias, journals, newspapers, transcripts, images, maps, and more. The National Geographic site contains many current video clips that have been indexed for Alberta Programs of Study. The content is organized by grade level, subject, and curriculum objective. Use the search engine to quickly find key concepts. Check this site often as new interactive multimedia segments are being added all the time.

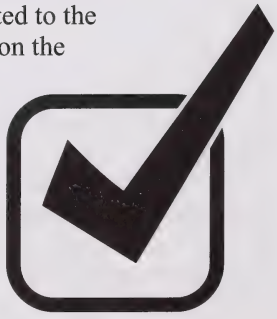
If you find a password is required, contact your teacher or school to get one. No fee is required.

Assessment and Feedback

Each Student Module Booklet is divided into sections. Within each section, your work is grouped into lessons. Within the lessons, there are readings, activities, investigations, and questions for you to do. By completing these lessons, you will discover scientific concepts and skills, practise or apply what you have learned, and develop a positive attitude toward science.

Suggested answers to the questions are provided at the end of each lesson. They provide you with immediate feedback to the questions you complete in the lessons.

At several points throughout the Student Module Booklet, you will be directed to the accompanying Assignment Booklets. Your grading in each module is based on the assignments you submit for assessment. You will be asked to submit each Assignment Booklet to your teacher once it is completed. Your teacher will then mark your Assignment Booklet and give you feedback as to how you are doing. Once you get your Assignment Booklet back, be sure you review your teacher's comments and correct any errors you made.

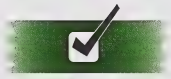


Special Features

The Student Module Booklets also contain some special features.

Going Further activities are provided throughout the course. These are optional activities for students who want to extend their knowledge about the material just covered.

Did You Know? boxes appear throughout the course. They provide additional, interesting, sometimes quirky, facts about the material just covered.



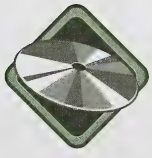
These parts of the Student Module Booklets contain the Glossary, Suggested Answers, and Image Credits for the lesson preceding them.

Visual Cues

You will find many visual cues throughout this course. One such cue is the use of colour to highlight the important terms that are defined in the Glossary at the end of each lesson. You will also encounter some icons in the margins. Read the following explanations to find out what each icon prompts you to do.

Refer to the textbook.

Use the Internet.



Refer to the multimedia CD given in Module 1.

Pay special attention to the safety precautions mentioned.

Answer questions in the Assignment Booklet.

Prepare for a lab that is to be completed at a later date.

Module Overview

Offshore drilling has increased significantly since 1990. Today, offshore drilling mainly consists of drilling for conventional oil and natural gas. In the future, offshore drilling may be used for obtaining methane hydrates beneath the ocean floor. Methane hydrate is a combination of ice and natural gas. It looks like ice, but it burns like a candle. The offshore deposits of methane hydrate that have been found are immense. Two relatively small deposits (not much smaller than Prince Edward Island) off the coasts of North Carolina and South Carolina are estimated to contain about 35 trillion cubic metres of methane. This is about 500 times Canada's natural-gas consumption in 2002. Methane hydrate deposits are a significant potential energy source for the future.

The burning of methane hydrate is just one example of the chemical changes and reactions you will study in this module. The burning of methane, however, results in the increased levels of carbon dioxide in the atmosphere, leading to increased global warming. The burning of hydrogen gas, on the other hand, results in the production of water vapour only and, thus, is much kinder to the atmosphere.

In this module you will study transformations of matter in chemical changes and how energy is involved. You will develop a basic understanding of the atomic structure of matter and how elements combine to form compounds. You will identify and classify chemical changes, name compounds, and write word equations and balanced formula equations to represent chemical reactions. You will also be introduced to the mole concept and the law of conservation of mass.

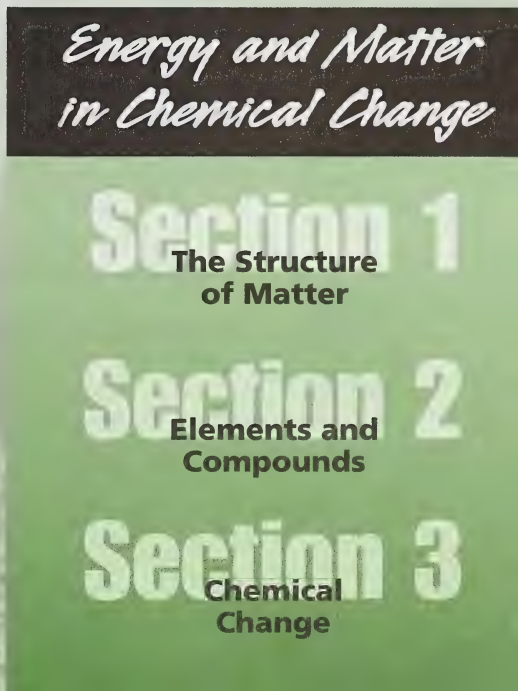
Turn to page 3 of the textbook and read "Focus on the Nature of Science" to see what is ahead in Unit A. Then read "Exploring" on pages 4 and 5 for more information on methane hydrate.

Assessment

This module, Energy and Matter in Chemical Change, has three section assignments. The mark distribution is as follows:

Assignment Booklet 1A	
Section 1 Assignment	32 marks
Assignment Booklet 1B	
Section 2 Assignment	66 marks
Assignment Booklet 1C	
Section 3 Assignment	42 marks
TOTAL	**140 marks**

Be sure to check with your teacher if this mark allocation is valid for you. Some teachers may include other reviews and assignments for additional assessment.



The Structure of Matter

What types of materials do you have in your home? Are the materials natural or produced through technologies? Over thousands of years, people have learned to work with the properties of different materials to produce the items they want.

Today, scientists work with the understanding that the clothes you wear, the food you eat, and even the water you drink are all made up of particles. The understanding that particles make up the underlying structure of matter has allowed scientists to discover new materials and invent various chemical technologies.

In this section you will review lab safety rules, WHMIS safety symbols, as well as safety hazard symbols on products found around the home. You will identify the differences between physical and chemical properties and learn to recognize several types of chemical reactions. You will then describe various ways in which early human societies discovered and used naturally occurring materials. You will finish this section by observing how the understanding of the composition of matter gradually developed over the past few centuries and how experimental evidence led to various models of the atom.

Turn to page 6 of the textbook and read the introduction to Unit A 1.0. Pay particular attention to the key concepts and the learning outcomes listed in the left margin. They provide a quick overview of what you will cover in this section.

Lesson 1

Safety in the Laboratory

Safety is an essential part in any laboratory, whether it is a scientist working with sophisticated equipment or you, a student chemist, working with basic lab equipment and chemicals. You must always be aware of the potential hazards of both the equipment and chemicals you are using. In addition to knowing the hazardous materials symbols, you need to observe a number of safety rules.

Turn to page 7 of the textbook and read the introductory paragraphs of "Safety in the Laboratory." Then carry on by reading "Understanding the Rules" and studying the Science Laboratory Safety Rules listed.

1. An important protection item worn in a science lab are _____.

2. Why should you wash your hands thoroughly after each activity, especially after handling chemicals?

3. Why should you tie back loose hair and roll up loose sleeves?

4. Why is it important to label any container you put chemicals into?

Check your answers with those on pages 21 and 22.

WHMIS: a system that provides details on how to store, handle, and dispose of any chemical found in the workplace

Safety hazard symbols are used to label containers of chemicals both in the workplace and in the home. The system used in the workplace is called **WHMIS**—**W**orkplace **H**azardous **M**aterials **I**nformation **S**ystem. Since a science laboratory in a school is considered a workplace, WHMIS is used.

There are eight WHMIS symbols altogether. Study these symbols and the short description for each.

HAZARD SYMBOLS

CLASS A

Compressed gas

CLASS B

Flammable and combustible material

CLASS C

Oxidizing material

CLASS D

Poisonous and infectious material

1. Materials causing immediate and serious toxic effects

CLASS D

Poisonous and infectious material

2. Materials causing other toxic effects

CLASS D

Poisonous and infectious material

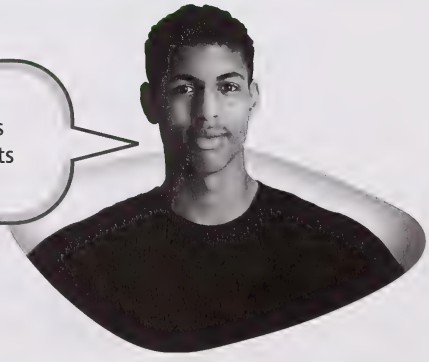
3. Biohazardous infectious material

CLASS E

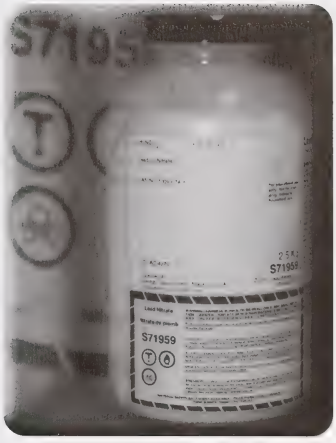
Corrosive material

CLASS F

Dangerously reactive material



Turn to page 8 of the textbook and read "WHMIS." Become familiar with the WHMIS symbols and descriptions for each type of hazardous substance. If you are in a classroom and have access to a supervised laboratory, you will see some of these symbols on the containers of chemicals you will use in this course.



Here are some examples of the various classes of chemicals you might encounter in a Science 10 laboratory or at home:

- **Class A: Compressed gas**

  natural gas lines

- **Class B: Flammable and combustible material**

  solids—calcium, fire starters, lithium batteries, magnesium, mothballs, sodium

  liquids—adhesives, camping fuel, aftershave, gasoline, gasoline additives, paint thinners, rubbing alcohol, solvents, turpentine

- **Class C: Oxidizing material**

  adhesives, ammonium nitrate fertilizer, bleaching powder, disinfectants, fiberglass repair kits, hair and textile dyes, hydrogen peroxide

- **Class D1: Poisonous and infectious material causing immediate and serious toxic effects**

  arsenic, chloroform, disinfectants, dyes, fungicides, gas additives, herbicides, insecticides, wood preservatives

- **Class D2: Poisonous and infectious material causing other toxic effects**

  acetone, asbestos, toluene, diisocyanate

- **Class E: Corrosive material**

  acids, caustic soda (lye or sodium hydroxide), disinfectants, drain cleaners, dyes, household bleach, oven cleaners, paint remover, rust remover

5. What does WHMIS stand for?
6. Which WHMIS symbols will you encounter in a school science laboratory?

Check your answers with those on page 22.

MSDS: an information sheet that identifies the chemical and the physical properties of a substance, safe handling, first aid, and cleanup procedures in case of spill or leak

Another part of WHMIS is the use of a **Material Safety Data Sheet (MSDS)** for every chemical used in the workplace. An MSDS provides detailed information about each chemical. The information includes physical data, such as melting point, boiling point, toxicity, health effects, first aid, and cleanup procedures in case of a spill or leak.

Read "Material Safety Data Sheets" on page 9 of the textbook. Closely study the sample MSDS in Figure A1.5.

7. What is the chemical name of the substance listed in the sample MSDS in Figure A1.5 closely.

8. What do the WHMIS symbols indicate about the product?

Check your answers with those on page 22.

Quicklab

Reading a Material Safety Data Sheet

Use the MSDS (Material Safety Data Sheet) at the back of this Student Module Booklet and the WHMIS label given to complete this activity.

9. a. What is the chemical name of this substance?

   b. How should this substance be stored?

   c. Is there a fire hazard when this substance is used?

   d. How should a spill be cleaned up?

   e. What are the medical effects of this substance?

Calcium Chloride Anhydrous
Chlorure de Calcium Anhydre

THIS IS NOT A CONTROLLED SUBSTANCE

Normal precautions should be taken to minimize any exposure to this or any chemical. Handling care generally in keeping with safe laboratory practice is recommended.

MATERIAL SAFETY DATA SHEET IS AVAILABLE

CECI N'EST PAS UN PRODUIT CONTROLÉ

Des précautions usuelles devraient être prise pour

   f. What is the recommended first aid if you are exposed to this substance?

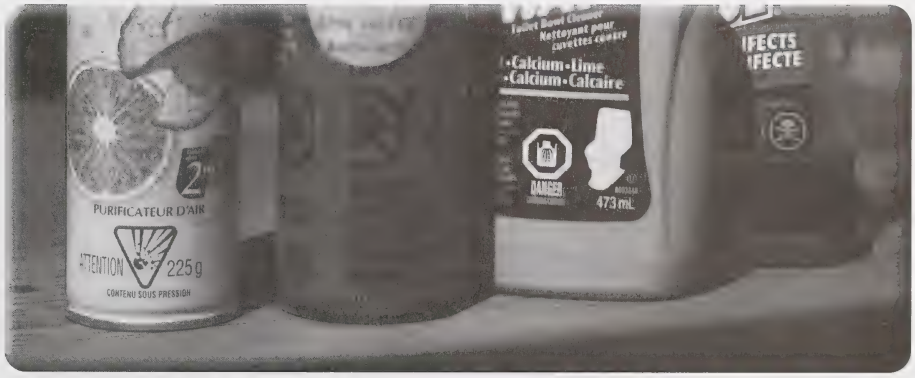
10. If using this product, what safety precautions would you take?

11. Which provides more information—the label or the MSDS?

12. Why isn't all the safety information provided on the product label?

Check your answers with those on pages 22 and 23.

Many products used in the home and around the yard are hazardous. Turn to page 8 of the textbook and read "Safety Hazard Symbols."



The information in the reading has since been updated by Health Canada. The information symbols (poisonous, flammable, etc.) shown in Figure A1.2 on page 8 of the textbook are still used. The three shapes (octagon, diamond, and triangle) for the degree of severity and the different colours, however, are no longer used. Presently, only two shapes are used: the octagon and the triangle. The octagon shows that the warning refers to the contents of the container, and the triangle shows that the warning refers to the container itself. Because some containers use a propellant that may explode when heated, a warning label is now given regarding the container. Note that the symbols in the preceding photo and on the Varathane can in Figure A1.3 on page 8 of the textbook follow the updated information system.

Visit the following Health Canada site for information on safety hazard symbols:

http://www.hc-sc.gc.ca/hecs-sesc/cps/publications/hazard.htm

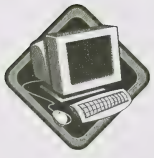
13. What do the safety hazard symbols on the product in the following photograph say about the product?

14. Draw the safety hazard symbol for each of the following situations.

    a. The contents of a product are corrosive.
    b. The contents of a product are flammable.
    c. The container of a product is explosive.

Check your answers with those on page 23.

Safety in the laboratory is an important part of any science course. The WHMIS symbols and the MSDS provide information on safety. Another important part of safety in the laboratory is how you work and behave in the laboratory. You need to know how to use safety equipment and how to respond if an emergency does occur. You may want to read "Student Reference 1: Safety" on pages 455 and 456 of the textbook from time to time to remind yourself of the WHMIS symbols and laboratory safety practices.



Visit the following website to test yourself on laboratory safety, WHMIS, and Material Safety Data Sheets.

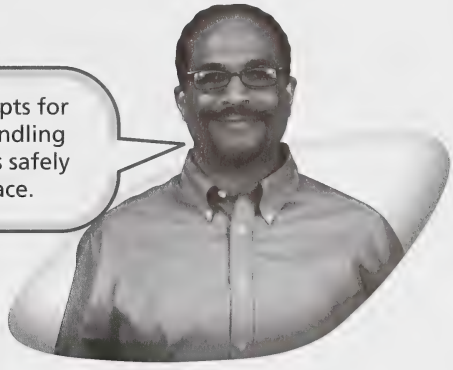
http://www.scienceman.com/science10

Once there, click on "Unit A: Tech Ideas" and scroll down to Text Pages 7–10. You will find a number of links to games that will test your knowledge.

Using and transporting chemicals presents another safety issue—environmental safety. Improper chemical disposal and chemical spills can seriously harm the environment. For this reason, there are many regulations governing the production, transport, storage, use, and disposal of chemicals. If you are using chemicals in a laboratory or at home, be aware of how to use a chemical and how to dispose the container. Read the labels on the containers carefully.

Now, turn to page 11 of the textbook and read "Environmental Safety."

Looking Back


You have just completed the concepts for this lesson. You discovered that handling chemicals and performing activities safely is very important in the workplace.

15. Answer questions 1, 2, 11.a., and 11.c. of "Check and Reflect" on page 11 of the textbook.

Check your answers with those on page 23.

Go to pages 1 and 2 of Assignment Booklet 1A and answer questions 1 to 4.

Glossary

Material Safety Data Sheet (MSDS): an information sheet that identifies the chemical and the physical properties of a substance, safe handling, first aid, and cleanup procedures in case of a spill or leak

WHMIS: Workplace Hazardous Materials Information System; a system that provides details on how to store, handle, and dispose of any chemical found in the workplace

Suggested Answers

1. An important protection item worn in a science lab are **safety glasses**.

2. You should wash your hands after each activity, especially after handling chemicals, in case you accidentally got some toxic chemical on your hands.

3. You should tie back loose hair and roll up loose sleeves so you don't accidentally contact the flame of a Bunsen burner with your hair or sleeve. Loose sleeves can also get caught on or in things, cause spills, and absorb chemicals or liquids.

4. You need to label any container you put chemicals into so you know exactly what is in each container. It is surprising how quickly one can forget or mix up two unlabelled containers.

5. **W**orkplace **H**azardous **M**aterials **I**nformation **S**ystem

6. The WHMIS symbols you may encounter in a school science laboratory are compressed gas (class A), flammable and combustible material (Class B), oxidizing material (Class C), poisonous and infectious material causing immediate and serious toxic effects (Class D1), poisonous and infectious material causing other toxic effects (Class D2), and corrosive material (Class E). You most likely won't see any biohazardous material (Class D3) or dangerously reactive material (Class F) in a school laboratory.

7. Sulfuric acid is the product listed in Figure A1.5.

8. The WHMIS symbols indicate that the substance is poisonous and corrosive.

9. a. The chemical is calcium chloride anhydrous (or simply calcium chloride).

   b. This substance should be stored in a tightly closed container in a cool, dry, well-ventilated area. Keep it away from flame and sparks. Exposure to atmosphere allows calcium chloride to absorb water and form a solution.

   c. This substance is not considered to be a fire or explosion hazard.

   d. The area of a spill must be ventilated. Spills are to be swept up and put into a container for reclamation or disposal. Use vacuuming or wet sweeping to avoid dust dispersal. Small amounts of residue may be flushed to the sewer with plenty of water. Avoid inhalation, ingestion, and contact with eyes or skin.

   e. Acute exposure causes moderate irritation to skin, eyes, and the respiratory tract. Chronic (long-term) exposure may produce "soda ulcers" on hands and perforation of the nasal system.

   f. If the substance contacts an individual's eyes, immediately wash eyes with plenty of water for at least 15 minutes, lifting upper and lower eyelids occasionally. Seek immediate medical attention.

      If the substance contacts skin, wipe off excess material and wash with plenty of soap and water for at least 15 minutes. Remove contaminated clothing and shoes. Seek medical attention if irritation occurs or persists.

      If the substance is inhaled, take the individual out for fresh air. If the individual is not breathing, give artificial respiration. If breathing is difficult, give oxygen. Seek medical attention.

      If the substance is ingested, give large quantities of water to dilute it. Induce vomiting immediately as directed by medical personnel. Never give anything by mouth to an unconscious person. Seek immediate medical attention.

10. If using this product, provide adequate ventilation and access to water and shower. Make sure you wear appropriate protective equipment, such as goggles and a lab coat.

11. The MSDS provides more information than the label.

12. All safety information is not provided on the label because there is not enough room.

13. The safety hazard symbols on the product show that the contents are poisonous and flammable and that the container is explosive.

14. a. b. c.

15. **Textbook questions 1, 2, 11.a., and 11.c. of "Check and Reflect," p. 11**

    1. It is important for all students to follow the safety rules in a science class so nobody gets injured.

    2. a. The following precautions minimize the risk of poisoning:

        - Never eat or drink anything in a science laboratory.
        - Use gloves when required, and never put your fingers in your mouth after handling chemicals or equipment.
        - Wash your hands thoroughly with soap and water once you are finished in the science laboratory.

       b. The following precautions minimize the risk of scalding:

        - Use beaker tongs when handling a hot container.
        - Boil small quantities of water.
        - Keep a safe distance away from a boiling container.

       c. The following precautions minimize the risk of eye damage:

        - Wear safety glasses whenever performing a lab or near someone performing a lab.
        - Avoid bringing your face close to containers into which you are pouring or heating chemicals.

    11. a. Flush the affected eye immediately with large quantities of water using either a prescribed washing station or a cold water faucet. If using a faucet, cup your hand and splash water into the affected eye. Also, get medical help immediately.

        c. Immediately turn off or unplug the heat source or shut off the gas valve for the bunsen burner. If the beaker of water is in danger of boiling over, remove it from the heat source and leave the building according to fire evacuation procedures.

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Lesson 2

Properties and Classification of Matter

While growing up, did you have a marble collection? If so, did you organize them according to colour, shape, or size? Did you use any other special characteristics to separate your marbles?

physical property: a property that describes the physical appearance and composition of a substance

chemical property: a property that describes the reactivity of a substance

Similarly, scientists use various characteristics to separate chemicals. **Physical properties**—such as boiling point, melting point, state, and colour—and **chemical properties**—such as ability to burn (combustion), reaction with water, and reaction with acids—are used to identify substances.

Turn to page 13 of the textbook and read "Properties Used to Classify Substances." Pay special attention to the information in Tables A1.1 and A1.2. Also, take note of the information given in Figures A1.7 and A1.8.

1. How do physical properties differ from chemical properties?

2. a. List three physical properties.
   b. List three chemical properties.

3. What is meant when a substance is said to be malleable?

4. What is the difference between the chemical properties ability to burn and flash point?

5. Give an example of the chemical property behaviour in air.

Check your answers with those on page 30.

pure substance: a substance in which all particles are identical

mixture: a combination of pure substances

element: a pure substance that cannot be broken down into other substances

compound: a chemical combination of two or more elements in a specific ratio

Physical and chemical properties are used to identify **pure substances**. However, not all substances are pure substances. Matter is divided into pure substances and **mixtures**. Each of these classes of substances is further subdivided. Pure substances are subdivided into **elements** and **compounds**; and mixtures are subdivided into solutions, mechanical mixtures, suspensions, and colloids. Study Figure 1.1 closely.

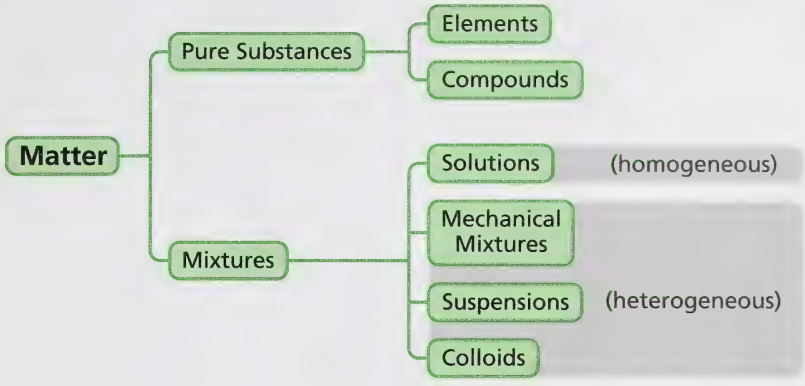


Figure 1.1: Classification of matter

In previous science courses you studied the classification of matter. For a review of the classes of matter, read "Pure Substances and Mixtures" on pages 14 and 15 of the textbook.

6. How are elements and compounds different?

7. How are mixtures different from pure substances?

8. What is a solution?

9. What is a mechanical mixture?

10. Give an example of a solution and an example of a mechanical mixture.

11. How is a homogeneous mixture different from a heterogeneous mixture?

Check your answers with those on pages 30 and 31.

In this lesson you identified some physical and chemical properties. Physical properties are fairly easy to see. Chemical properties are identified through a chemical change. A chemical change is what chemists call a **chemical reaction**. The various chemical properties can be observed directly or indirectly during chemical reactions.

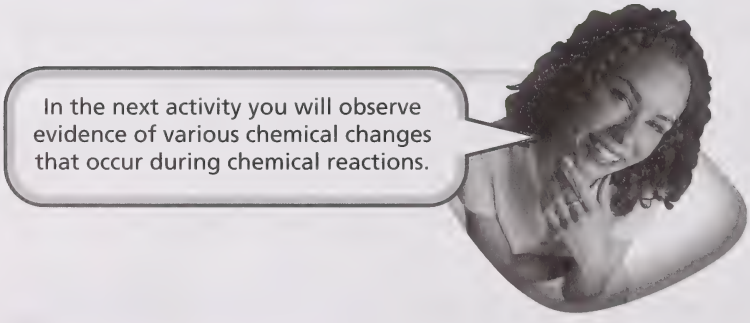
chemical reaction: a process that occurs when substances react to form a new substance or substances

Turn to page 15 of the textbook and read "Chemical Reactions."

12. What is a chemical reaction?

13. Give two examples of chemical properties that can be observed during a chemical reaction.

Check your answers with those on page 31.



Inquiry Lab

Evidence of Chemical Change

Turn to page 16 of the textbook and read the entire activity.

If you have access to a supervised laboratory, do **Part A**. If you do not have access to a supervised laboratory, do **Part B**.

Part A

Follow the procedure as described on page 16 of the textbook. **Pay particular attention to the safety precautions mentioned.**

14. Copy and complete a table similar to the following.

Reaction	Observations	
	Before Reaction	During and After Reaction
zinc metal and hydrochloric acid		
sodium hydroxide solution and phenolphthalein indicator		
magnesium metal strip and dilute sulfuric acid		
hydrochloric acid and bromothymol blue indicator		
sodium chloride solution and silver nitrate solution		
iron(III) chloride solution and sodium hydroxide solution		
magnesium strip metal and vinegar		
candle wax and oxygen		

15. Answer the following on page 16 of the textbook.

    a. questions 1, 2, and 3 of "Analyzing and Interpreting"
    b. question 4 of "Forming Conclusions"

Check your answers with those on pages 31 and 32.

Part B

Insert the Science 10 Multimedia CD into your computer. View the segments "Investigating Chemical Reactions: Evidence of Chemical Change, Part 1" and "Investigating Chemical Reactions: Evidence of Chemical Change, Part 2." You will watch a number of investigations that show evidence of chemical changes.

16. Copy and complete the following table by viewing the results of Work Stations 1, 3, 4, and 5 in the segments.

Reaction	Observations	
	Before Reaction	During and After Reaction
toilet bowl cleaner and water		
hydrochloric acid and bromothymol blue and antacid tablet		
calcium chloride solution and sodium carbonate solution		
hydrochloric acid and mossy zinc metal		

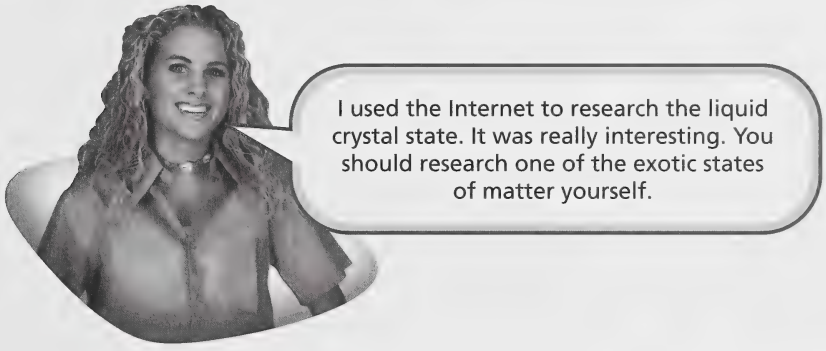
17. List the observations you made in this activity that indicated that a chemical reaction had taken place.

Check your answers with those on pages 32 and 33.

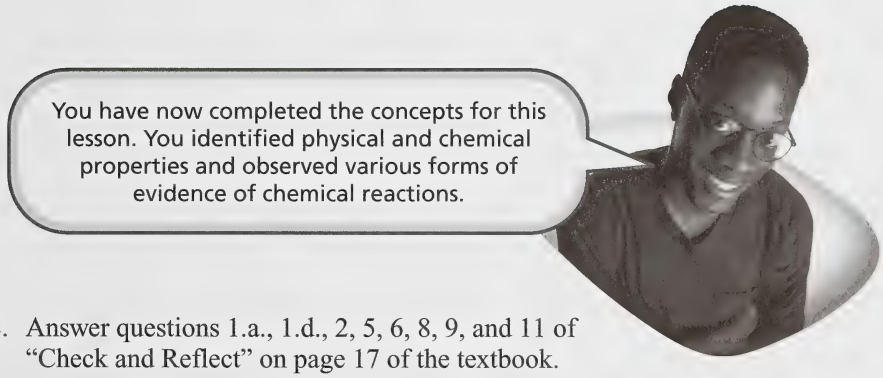
Writing lab reports is a part of every science course. You will be asked to write a lab report for an activity later in this module. Since much of the report is already presented in the textbook lab activity and in the observations you record as you do the investigation, you will not be required to do a formal lab report for every activity.

You have just witnessed several types of evidence of chemical reactions. In science you need to be able to recognize chemical reactions. All reactions share certain characteristics.

Turn to page 17 of the textbook and read "Recognizing Chemical Reactions." Then read the information given in "*re*SEARCH."


I used the Internet to research the liquid crystal state. It was really interesting. You should research one of the exotic states of matter yourself.

Looking Back


You have now completed the concepts for this lesson. You identified physical and chemical properties and observed various forms of evidence of chemical reactions.

18. Answer questions 1.a., 1.d., 2, 5, 6, 8, 9, and 11 of "Check and Reflect" on page 17 of the textbook.

Check your answers with those on page 33.

Go to pages 3 and 4 of Assignment Booklet 1A and answer questions 5 to 13.

Glossary

boiling point: the temperature at which a substance at normal air pressure changes from a liquid to a gas

chemical property: a property that describes the reactivity of a substance

chemical reaction: a process that occurs when substances react to form a new substance or substances

melting point: the temperature at which a substance at normal air pressure changes from a solid to a liquid

mixture: a combination of pure substances

physical property: a property that describes the physical appearance and composition of a substance

pure substance: a substance in which all particles are identical

solubility: the ability of a substance to dissolve in another substance

Suggested Answers

1. Physical properties describe the physical appearance of a substance and can be observed without affecting its chemical composition. Chemical properties describe the reactivity of a substance.

2. a. Any three properties listed in Table A1.1 on page 13 of the textbook may be listed.
   b. Any three properties listed in Table A1.2 on page 13 of the textbook may be listed.

3. A substance is said to be malleable when it can be beaten or rolled into sheets without it crumbling or breaking apart.

4. The ability to burn means that the substance can support combustion—produce a flame, heat, and light. Flash point is the lowest temperature at which a substance will ignite when a flame is applied.

5. Some examples include iron corroding, copper turning green, and aluminium turning black.

6. Elements are made up of only one type of atom and cannot be broken down into simpler substances. Compounds contain two or more elements in a specific ratio. Compounds can be separated into the individual elements that make up the compound.

7. In a pure substance, all the particles are identical. A mixture is made up of a number of pure substances, thus containing more than one type of particle.

8. A solution is a type of mixture in which the different types of particles are not visible.

9. A mechanical mixture is a mixture in which the different types of particles remain separate and are visible. A mechanical mixture is one type of heterogeneous mixture.

10. Answers will vary. Sugar dissolved in water is an example of a solution. Soil is an example of a mechanical mixture.

11. A homogeneous mixture looks the same throughout, whereas the different types of particles in a heterogeneous mixture are visible.

12. A chemical reaction is a process that occurs when substances react to form a new substance or substances.

13. Answers will vary. Two examples of chemical properties that can be observed during a chemical reaction are

    - hydrogen combining with oxygen to form water
    - an iron nail rusting

14. Your completed table should be similar to the following.

Reaction	Observations	
	Before Reaction	During and After Reaction
zinc metal and hydrochloric acid	• shiny, solid zinc • colourless aqueous solution	Bubbles appear and some zinc dissolves.
sodium hydroxide solution and phenolphthalein indicator	• colourless aqueous solution	Aqueous solution turns pink when the indicator is added to sodium hydroxide.
magnesium metal strip and dilute sulfuric acid	• shiny magnesium metal • colourless aqueous solution	Bubbles appear and some magnesium dissolves.
hydrochloric acid and bromothymol blue indicator	• colourless aqueous acid • aqueous blue indicator solution	Aqueous blue solution turns green and then yellow when the indicator is added.
sodium chloride solution and silver nitrate solution	• colourless aqueous solutions	Mixture turns cloudy white after the solutions are mixed.
iron(III) chloride solution and sodium hydroxide solution	• yellow aqueous solution • colourless sodium hydroxide solution	Brown (rust-coloured) precipitate results after the solutions are mixed.
magnesium strip metal and vinegar	• shiny magnesium metal • colourless aqueous acid	Bubbles appear and the magnesium dissolves.
candle wax and oxygen	• solid wax • colourless air	Heat and light are released, and some wax melts and disappears.

15. a. **Textbook questions 1, 2, and 3 of "Analyzing and Interpreting," p. 16**

1. In each case where a metal was mixed with an acid, bubbles formed and some of the metal dissolves.

2. In each case where a solid precipitate or cloudiness was produced, the reactants were ionic compounds. Recall from your previous science courses that an ionic compound results when electrons are transferred during a chemical reaction between a metal and a non-metal or a polyatomic ion.

3. The evidence that a burning candle involves a chemical reaction is that energy is released in the form of heat and light and that the wax disappears over time.

b. **Textbook question 4 of "Forming Conclusions," p. 16**

4. The types of observations made that indicated a chemical reaction was taking place included

- the appearance of bubbles
- the dissolving of a metal
- the formation of a solid precipitate or cloudiness
- a colour change
- a temperature change
- a release of energy in the form of heat or light

16. Your completed table should be similar to the following.

Reaction	Observations	
	Before Reaction	During and After Reaction
toilet bowl cleaner and water	• toilet bowl cleaner (in crystal form) • water	The crystals dissolve, the solution turns blue, bubbles appear, and the temperature rises from 21°C to 23°C.
hydrochloric acid and bromothymol blue and antacid tablet	• colourless aqueous hydrochloric acid solution • yellow solution of bromothymol blue and antacid tablet	Bubbles appear and the yellow colour disappears from the solution.
calcium chloride solution and sodium carbonate solution	• colourless aqueous solution • heat released when calcium chloride crystals were added to water • colourless aqueous sodium chloride solution	Bubbles appear and the solution turns cloudy but clears after the bubbles escape.
hydrochloric acid and mossy zinc metal	• colourless aqueous hydrochloric acid solution • dull grey metal	Bubbles appear and the surface of the mossy zinc metal dissolves and becomes shiny.

17. The observations that indicated that a chemical reaction was taking place included

    - the appearance of of bubbles
    - the dissolving of the surface of a metal and surface becoming shiny
    - the formation of a solid precipitate or cloudiness
    - a colour change
    - a temperature change

18. **Textbook questions 1.a., 1.d., 2, 5, 6, 8, 9, and 11 of "Check and Reflect," p. 17**

    1. a. *Boiling point* is the temperature at which a substance changes from a liquid to a gas.
       d. *Solubility* is the ability of a substance to dissolve in another substance.

    2. a. Two clear solutions are added together and the resulting mixture becomes cloudy or a new substance settles at the bottom of the liquid.

       b. Bubbles appear in a liquid that is not at its boiling point.

       c. A thermometer, inserted in the substance, shows an increase in temperature; or the container becomes warm to touch.

    5. All chemical reactions produce new substances with new properties, and all chemical reactions involve a flow of energy.

    6. Because the substance melts over a range of temperatures, it is not a pure substance. Pure substances melt at specific temperature.

    8. This material should be classified as a mechanical mixture.

    9. Homogenized milk is a heterogeneous substance because it is still made up of small droplets of cream and the watery substance. It just looks like it is homogeneous because the cream droplets are so small and evenly dispersed.

    11. If you know the boiling point of the liquid, you can determine if the bubbles are from boiling or from a chemical reaction. If the bubbles are occurring at normal pressures and at a temperature below the boiling point of the liquid, they must be from a chemical reaction. You can also collect the gas released and try to condense it into a liquid. If it is still a gas at normal pressure and at a temperature below the boiling point of the liquid, the gas must be a new substance produced by a chemical reaction.

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Lesson 3

Developing Ideas About Matter

Humans have worked with chemicals and chemical processes for thousands of years. The Ancient Egyptians had cosmetics; most First Nations people used several methods to preserve food and various plants for medicines; and the Inuit used copper for harpoons, arrows, knives, handles for pots, and so on. Each group had to discover how to work with the substance or process to make the substance more useable. In more recent years, ideas of what makes up matter developed and models, called **atoms**, to show the different properties of each element were created.

atom: the smallest particle of an element that has the properties of that element

To learn more about some of the processes people centuries ago became familiar with, read the introductory information of "Developing Ideas About Matter" and the information in "Food Chemistry" on pages 18 and 19 of the textbook.

1. Name two plants that were used for medicines by ancestors of the First Nations people.

2. Describe six methods of preserving food.

3. How does salting preserve food?

4. What benefit of fermentation did early explorers take advantage of?

Check your answers with those on page 40.

In addition to ideas about chemical processes in food preservation, people centuries ago knew about chemical processes in metallurgy. Inuit people along the Coppermine River used naturally occurring copper to make hunting equipment, much like the knife in Figure 1.1, as well as some items for the home. First Nations people in British Columbia and Alaska used copper metal to make ceremonial plaques. In both cases, the people learned that heating the copper metal allowed it to be hammered into shapes without the copper breaking up as it did when no heat was used.

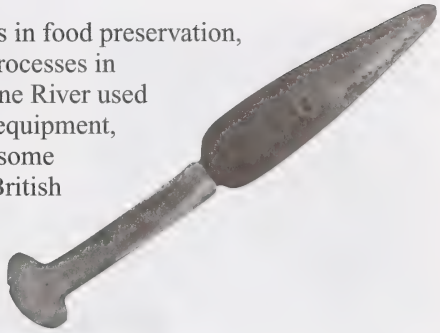

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Figure 1.2: Knife made from local copper and antler

Read "Metallurgy—An Early Branch of Chemistry" on pages 19 and 20 of the textbook.

5. The Sumerians, in the Middle East, began smelting tin ore with copper ore to form bronze. What benefit did bronze have over using natural copper?

Check your answers with those on page 40.

Did You Know?

An ancient copper mine in Mamainse Point, Ontario, was mined by First Nations people 6000 years before the time of the Egyptian pharaohs.

As civilizations advanced, some people began thinking about what made up matter. In particular, Greek philosophers, like Aristotle and Democritus, had different ideas on what made up matter.

Read the information in "Aristotle's Description of Matter" and "Alchemy" on pages 20 and 21 of the textbook.

6. What did the Greek philosopher Democritus propose about what made up matter?

7. How did Democritus's ideas about matter compare to those of Aristotle?

8. Why was Aristotle's ideas accepted over those of Democritus for so long?

9. What benefit did alchemists provide for chemistry?

Check your answers with those on page 40.

The scientific process that involves experimentation, observation, and forming conclusions forms the basis of the inquiry process in science. You are already familiar with much of the inquiry process as part of the lab reports you have completed in previous science courses.

Turn to pages 457 to 460 of the textbook and read "Student Reference 2: The Inquiry Process." This reading will provide you with a wealth of information for you to refer to while doing lab activities and writing lab reports.

The last five hundred years has brought about many new ideas and developments regarding the composition of matter. People like Robert Boyle (1627–1691), Antoine Lavoisier (1743–1794), John Dalton (1766–1844), J. J. Thomson (1856–1940), Ernest Rutherford (1871–1937), and Neils Bohr (1885–1962) have contributed evidence that has led to the development of a model of the atom.

Read the following information regarding the development of the model of the atom on pages 21 to 24 of the textbook.

- "Developing Hypotheses About Matter," page 21
- "John Dalton," page 22
- "J. J. Thomson," pages 22 and 23
- "Ernest Rutherford," page 23
- "Neils Bohr," page 24
- "The Quantum Mechanical Model of the Atom," page 25

10. What did Robert Boyle conclude based on his measurements of the relationships between volume and pressure of gases?

11. Antoine Lavoisier is credited with formulating the law of conservation of mass. What does the law of conservation of mass state?

12. John Dalton imagined that atoms were like small spheres. According to Dalton, how did elements differ in terms of these spheres?

13. a. How did J. J. Thomson's experiments dispute John Dalton's model of the atom?

    b. What are the particles that John Dalton's vacuum tube produced called today?

    c. What type of model of the atom did J. J. Thomson propose?

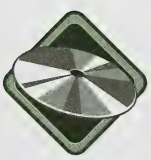
14. a. How did Ernest Rutherford's experiments refute J. J. Thomson's model of the atom?

    b. How is Ernest Rutherford's model of the atom different from the model proposed by J. J. Thomson?

15. a. Both Ernest Rutherford and Neils Bohr proposed a model of the atom that had negatively charged particles around a positively charged nucleus. How is the model of the atom proposed by Neils Bohr different from the model proposed by Ernest Rutherford?

    b. How did Neils Bohr know that negatively charged particles could move from one energy level to another?

Check Check your answers with those on pages 40 and 41.



Insert the Science 10 Multimedia CD into your computer, and view the segment "Rutherford Experiment." To start this segment, click on any of the angle labels.

16. Describe what happens when you click each angle label and the "All" label. What can you conclude from these observations?

Check Check your answer with the one on page 41.

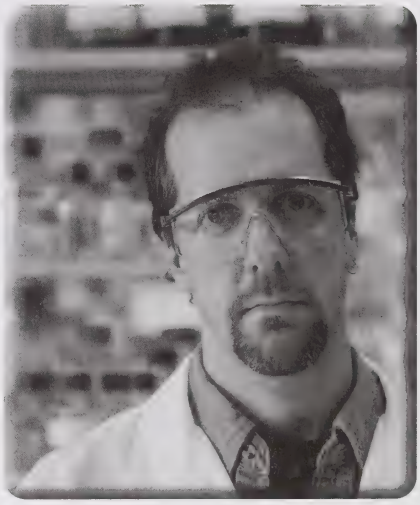
Today's model of the atom is based on a theory called quantum mechanics. This model is discussed briefly at the top of page 25 of the textbook. The discussion in the textbook suggests that the individual electron is a cloud of negative charge. Actually, the theory is that the electrons form a diffuse cloud of negative charge as they move about the nucleus of the atom.

This course, however, will use the particle view of an electron because it makes the chemistry easier to understand. On page 32 of the textbook, the electron is defined as a negatively charged particle that occupies different energy levels. This view of the electron helps to understand how atoms combine by transferring or sharing electrons.

17. How does the quantum mechanical model of the atom used today differ from the model of the atom proposed by Neils Bohr?

Check your answer with the one on page 41.

The field of chemistry provides many careers, from lab technicians to chemists to pharmacists to chemical engineers. Turn to page 26 of the textbook and read "Career and Profile." You will be introduced to Pauline Lee, a chemical engineer who works at Celanese Canada Inc. in Edmonton, Alberta.



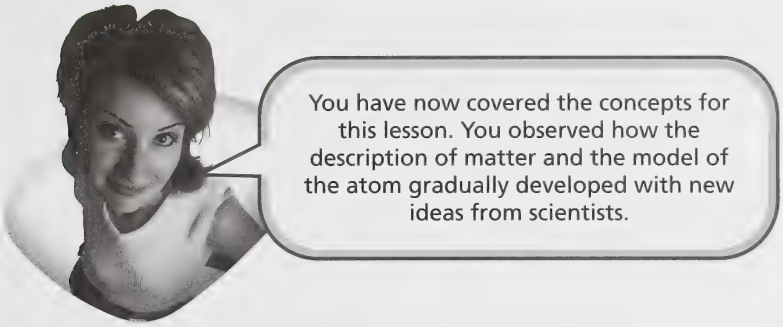
18. What types of jobs might a chemical engineer perform?

19. According to Pauline Lee, how is what you need to do as a process engineer different from what you do when studying in school?

Check your answers with those on page 41.

You can also look at various careers in chemistry on the Internet by entering the term *chemistry careers* into any of the Internet's search engines.

Looking Back


You have now covered the concepts for this lesson. You observed how the description of matter and the model of the atom gradually developed with new ideas from scientists.

20. Answer questions 1, 2, 5, 8, 9, and 10 of "Check and Reflect" on page 25 of the textbook.

Check your answers with those on pages 41 and 42.

Go to pages 4 and 5 of Assignment Booklet 1A and answer questions 14 to 19.

Glossary

atom: the smallest particle of an element that has the properties of that element

electron: a negatively charged particle in an atom that occupies energy levels around the nucleus

molecule: two or more atoms of the same element or of different elements bound together by covalent bonds

neutron: a neutral particle in the nucleus of an atom

nucleus: the positively charged centre of an atom made up of protons and neutrons, except hydrogen which has only a proton in its nucleus

proton: a positively charged particle in the nucleus of an atom

Suggested Answers

1. Ancestors of the First Nations people used the prickly rose and old man's whiskers for medicines.

2. Six methods of preserving food are drying, heating, freezing, fermentation, smoking, and chemical preservation.

3. Salting preserves food by drawing the moisture out of the food and out of any bacteria present, killing the bacteria or making them inactive.

4. Early explorers took advantage of the high vitamin C content in fermented (pickled) cabbage.

5. Bronze flowed more easily and was much stronger for tools and weapons.

6. Democritus proposed that matter was made up of tiny particles that could not be subdivided into smaller pieces.

7. Democritus's ideas were much more similar to today's ideas of what makes up matter than those of Aristotle.

8. Aristotle's ideas were accepted for so long because he was better known and well respected at the time he proposed his ideas.

9. Alchemists discovered new substances and procedures. They also improved laboratory equipment, such as glassware and distillation equipment.

10. Robert Boyle concluded that gases are made up of tiny particles that group together to make different substances.

11. The law of conservation of mass states that mass is neither produced nor lost during a chemical reaction.

12. John Dalton imagined that different elements were spheres of different size, mass, and/or colour.

13. a. J. J. Thomson's experiments showed that different elements all produced the same kind of beam. This suggested that different elements emitted particles that were identical.

    b. The particles are electrons.

    c. J. J. Thomson proposed a model of the atom in which negative particles were embedded in a positively charged sphere.

14. a. Ernest Rutherford theorized that according to J. J. Thomson's model, high-speed, positively charged particles should pass right through a sheet of gold foil. Rutherford's experiments, however, showed that some positively charged particles were deflected or bounced back when fired at a sheet of gold foil.

    b. Ernest Rutherford's model put the positive charge in the centre and had negatively charged particles moving around the positively charged centre. J. J. Thomson's model had the negative charges embedded in a positive sphere.

15. a. The model proposed by Neils Bohr had specific energy levels at which the negatively charged particles could exist.

    b. Neils Bohr knew that electrons could move from a higher energy level to a lower energy level because a certain colour of light was released.

16. **0°:** The particles go straight through.
    45°: The particles deflect 45° from the original path.
    90°: The particles deflect 90° from the original path.
    170°: The particles deflect 170° from the original path.
    All: Some of the particles deflect at various angles, and some particles go straight through.

    You can conclude that some particles are deflected at various angles by a sheet of gold foil.

17. In today's quantum mechanical model of the atom, electron movement is thought of as a "cloud" of negative charge. The electron "cloud" occupies the whole space of the atom at the different energy levels.

18. A chemical engineer might work in recovering and recycling materials, developing pharmaceuticals, purifying water, or refining gasoline.

19. According to Pauline Lee, you are given problems in school that you have to solve. As a process engineer, you need to find out what the problem is and then solve it.

20. a. **Textbook questions 1, 2, 5, 8, 9, and 10 of "Check and Reflect," p. 25**

    1. Methods of preserving food include drying, heating, freezing, salting, smoking, and fermentation. Drying, heating, freezing, and smoking are physical methods; salting is both a physical and chemical method; and fermentation is a chemical method.

    2. Iron was thought to be the rarest metal because it is not present in the metallic state except when found in a meteorite. Iron normally occurs on Earth as an ore that must be smelted to obtain metallic iron.

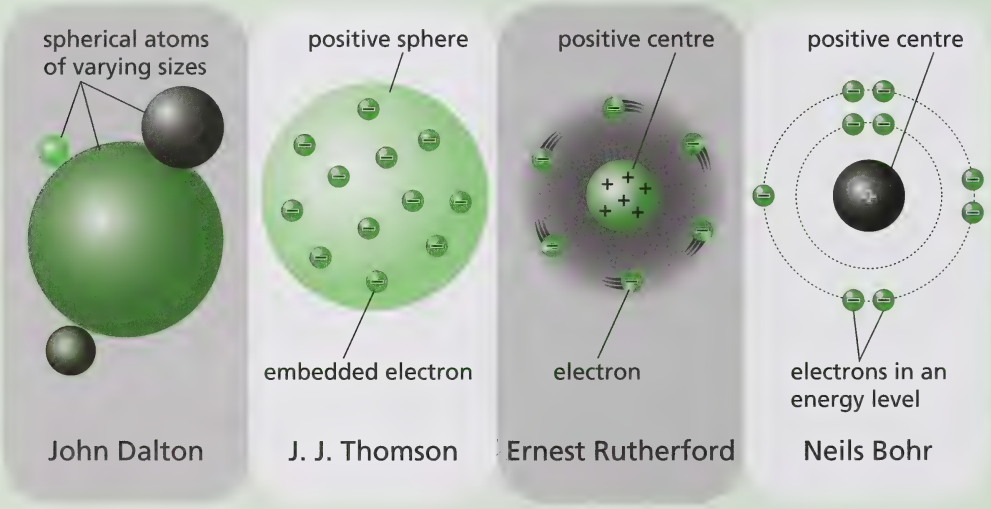
5. The four basic ideas proposed by John Dalton in describing the nature of matter are as follows:

    - All matter is made up of small, indivisible particles called atoms.
    - All the atoms of an element are identical in properties, such as size and mass.
    - Atoms of different elements have different properties.
    - Atoms of different elements can combine in specific fixed ratios to form new substances.

8. Neils Bohr examined the light released by hydrogen atoms when they are made to glow in a tube. He found that a particular colour of light was emitted when an electron went from a higher energy level to a lower energy level. When Bohr excited the electron to different energy levels, he found it emitted different colours of light at different energy levels.

9. The First Nations people of North America were able to find enough naturally occurring, metallic copper for their use. There was no need for them to develop the smelting technology.

10. Your diagrams should be similar to the following.



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Conclusion

In this section you reviewed safety rules in a laboratory, WHMIS symbols, as well as safety hazard symbols used on products for the home. You identified the differences between physical and chemical properties and learned to recognize several types of chemical reactions. You also described various ways in which early human societies discovered and used naturally occurring materials. You observed how the understanding of the composition of matter gradually developed over the centuries and how experimental evidence led to various models of the atom.

You now know that scientists theorize that everything around you is made up of particles and that this understanding of the structure of matter has led scientists to create many new materials. The creation of new materials requires a respect for chemicals and the hazards that these new materials may present.

Elements and Compounds

How many of the items shown in the photograph are familiar to you? Do you know which substances make up these different items?

You are most likely familiar with the periodic table of elements from previous science courses. Elements combine to form many substances, each with its own set of properties. It is these different properties of substances that provide all the different items people enjoy in today's industrial society.

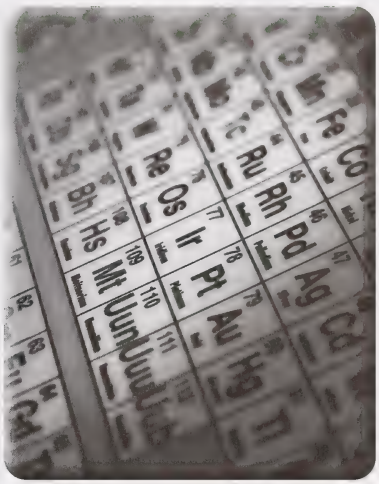
In this section you will review the periodic table and the elements that make up the table. You will describe the structure of atoms that make up the different elements, and you will develop an understanding of how atoms combine chemically by either gaining, losing, or sharing electrons. You will then become familiar with naming substances and categorizing them based on the properties they exhibit. To wrap things up, you will consider how toxic and hazardous chemicals affect people and the environment.

Turn to page 28 of the textbook and read the Unit A 2.0 introduction. Pay particular attention to the key concepts and the learning outcomes listed. They provide a brief overview of what you will cover in this section.

Lesson 1

The Periodic Table and Atomic Structure

In 1869, Dmitri Mendeleev created a periodic table based on the atomic weights of the known elements. He re-arranged the elements where their properties dictated. Mendeleev left gaps in the table for elements—unknown at the time—he predicted would be discovered later. These gaps were filled over the years with the names of elements that were predicted from the properties of elements already in the table.



The modern periodic table contains 114 elements. These 114 elements combine in various ways to form all substances known in the science community.

Elements and the Periodic Table

Turn to page 30 of the textbook and study the periodic table of elements. Notice that the elements are grouped into three differently coloured blocks. The pale green block of elements are **metals**; the orange block of elements are **non-metals**; and the purple block of elements are **metalloids**.

metal: a shiny, malleable, and ductile element

non-metal: an element with varying properties that are completely different from metals

metalloid: an element with properties intermediate between metals and non-metals

Now, turn to page 29 of the textbook and read "The Elements" for information on the three classes of elements. Read "*info*BIT" as well for an interesting tidbit.

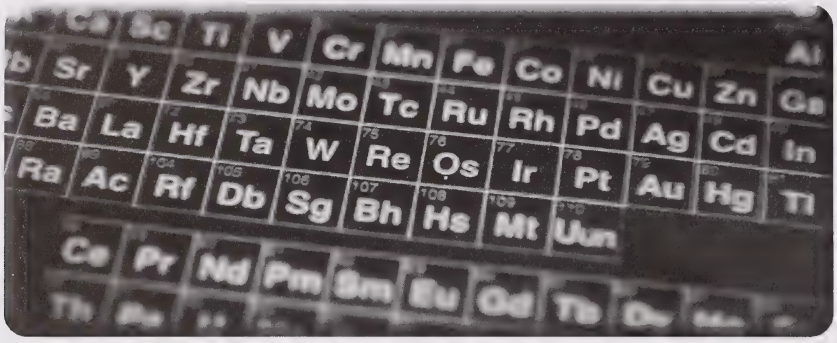
1. Name and describe three common properties of metals.
2. Describe three properties that set some metals apart from the other metals.
3. What characteristic do non-metals have in common?
4. Describe three properties that set some non-metals apart from the other non-metals.
5. What are metalloids?

Check your answers with those on page 54.

period: a horizontal row in the periodic table

group: a vertical column in the periodic table; also called family

In addition to the three classes, elements in the periodic table are further organized into rows and columns. Each row is called a **period**, and each column is called a **group**. All elements in the same row have the same number of energy levels. The elements in a group have the same number of electrons in their highest energy level and have similar chemical and physical properties.


Sc Ti V Cr Mn Fe Co Ni Cu Zn
Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In
Ba La Hf Ta W Re Os Ir Pt Au Hg Tl
Ra Ac Rf Db Sg Bh Hs Mt Uun
Ce Pr Nd Pm Sm Eu Gd Tb

Turn to page 31 of the textbook and read "The Periodic Table."

6. a. What are some common properties among most of the group 1 elements?
   b. What is the special name given to group 1 elements?
   c. Is hydrogen the same as the other group 1 elements?

7. a. What special name is given to the group 2 elements?
   b. How are group 2 elements similar to group 1 elements?

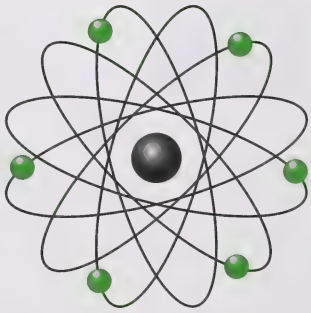
8. a. What special name is given to the elements in group 18?
   b. What is a common property of group 18 elements?

9. a. What special name is given to group 17 elements?
   b. What two common properties do group 17 elements possess?

Check your answers with those on page 55.

Atomic Theory

Elements are made up of particles called atoms. Atoms, in turn, are made up of three kinds of subatomic particles: electrons, protons, and neutrons. The study of how the subatomic particles are arranged in different atoms and what properties result from the various arrangements is called **atomic theory**.



Turn to pages 32 and 33 of the textbook and read the information in "Atomic Theory."

10. What is the smallest part of an element that still has the properties of that element?

11. State the charge of each subatomic particle that makes up an atom.

12. Where is most of the mass of an atom located?

13. What is an energy level?

14. Electrons with the greatest amount of energy are found in the energy level _____ away from the nucleus. Electrons with the _____ amount of energy are found in the energy level closest to the nucleus.

15. What is the maximum number of electrons at each level for the first three levels?

Check your answers with those on page 55.

Figure 2.1: Magnesium

All atoms of an element have the same number of protons and electrons. Atoms of the same element can contain different numbers of neutrons. The atomic number of an element is the number of protons in an atom. In the periodic table, the atomic number is given in the upper left of each square (as shown in Figure 2.1). The elements in the periodic table are listed in order of increasing atomic number.

16. Give the name of the element for each of the following atomic numbers, and state the number of protons in each element.

    a. 2 b. 8 c. 26

Check your answers with those on page 55.

Atoms of elements are also given a mass number and an **atomic molar mass**. To learn more, read "Mass Number and Atomic Molar Mass" on pages 33 and 34 of the textbook.

atomic molar mass: the average molar mass of the atoms of an element including all isotopes

17. What is the mass number of an element?

18. There is only one atomic number for each element. Why is there more than one atomic mass for each element?

Figure 2.2: Magnesium

Check your answers with those on page 56.

isotope: an atom of the same element containing a different number of neutrons

mass number: an integer equal to the total number of protons and neutrons in the nucleus of the atom

Atoms of the same elements with different numbers of neutrons are called **isotopes**. You may be familiar with the term *isotopes* with regard to isotopes of radioactive elements. Radioactive elements are not the only elements that have isotopes. Many common elements have isotopes as well. Each isotope is assigned a **mass number**.

19. Describe the three isotopes of hydrogen.

20. The symbol for a particular isotope of oxygen is ${}^{17}_{8}\mathrm{O}$. How many neutrons does this isotope of oxygen have?

21. a. What is atomic molar mass?
    b. What is the atomic molar mass of hydrogen?

Check your answers with those on page 56.

Formation of Ions

ion: an electrically charged atom or group of atoms

cation: a positively charged ion

anion: a negatively charged ion

Many elements form ions when they combine with other elements to form compounds. An **ion** is an electrically charged atom or group of atoms. Ions form when an atom gains or loses one or more of the electrons in their outer energy level. Positively charged ions are called **cations**. They form when an atom loses an electron. Negatively charged ions are called **anions.** They form when an atom gains an electron.

Figure 2.3: Sodium ion, Na^{1+}

Figure 2.4: Nitrogen ion, N^{3-}

Turn to pages 34 and 35 of the textbook and read "Formation of Ions".

22. Is the magnesium ion shown in Figure A2.9 a cation or an anion?

23. A sodium ion is written as Na^{+}, and a magnesium ion is written as Mg^{2+}. Why does sodium have a charge of 1+ and magnesium a charge of 2+?

24. Draw an element that has seven protons, eight neutrons, and seven electrons. Use the periodic table to identify the element.

25. An ion has 13 protons and 10 electrons. What is charge of the ion?

26. Which type of elements loses electrons? Which type gains electrons?

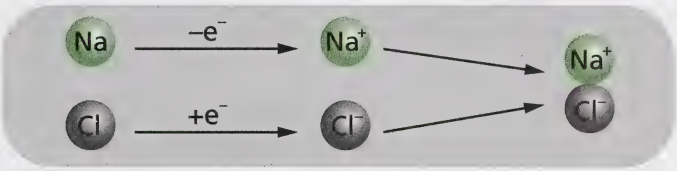
27. Why do atoms tend to gain or lose electrons?

Check your answers with those on page 56.

Elements Combining to Form Compounds

Certain elements combine to form compounds. Elements that combine to form these compounds are those that are reactive. Reactivity of an element is determined by the number of electrons the element has in its outer energy level. When the outer energy level is filled, the element is more stable. When the outer energy level is not filled, the element is more reactive. Electrons that are in the outer energy level are called **valence electrons**. Valence is the capacity of an atom to gain or lose electrons.

valence electron: an electron in the outer energy level of an atom



Turn to page 36 of the textbook and read "Elements Combine to Form Compounds."

28. Define *valence number*.

The partial periodic table given on page 36 of the textbook shows the electron arrangements of some of the elements in the first three periods (rows). By studying this table, you will find a number of interesting patterns.

29. Count the number of valence electrons in each element in each column in the periodic table.

    a. Complete the following table for each group or family shown in Figure A2.12 on page 36 of the textbook.

Group or Family	Number of Valence Electrons
1	
2	
13	

b. What did you notice about the number of electrons in the outer energy level for the elements in each group or family?

c. How do the properties of the elements of a group or family compare?

d. How do the number of valence electrons change as you move from left to right in the periodic table?

e. How does the number of energy levels change from one period to the next?

Check your answers with those on pages 56 and 57.

octet rule: a rule stating that atoms bond in such a way that each valence energy level has eight electrons

A good way to remember whether atoms gain or lose electrons when they form ions is to look at their energy levels. Atoms gain or lose electrons so they end up with a completely filled outer energy level. Many atoms gain or lose electrons to end up with eight electrons in their outer energy level. This has resulted in a rule called the **octet rule**. However, atoms that need only two electrons in their outer energy level are the exception.

Turn to page 38 of the textbook and read "The Octet Rule." Study the atoms in Figure A2.13 closely.

30. In Figure A2.13, state what is required for each atom in the diagram to have eight electrons in its outer energy level.

31. Why is it more difficult for metals to lose more than about three electrons.

Check your answers with those on page 57.

Problem-Solving Investigation

Classifying Unknown Liquids

Read the entire activity on page 37 of the textbook.

If you have access to a supervised laboratory, do **Part A**. If you do not have access to a supervised laboratory, do **Part B**.

Part A

Materials and Equipment

Along with the materials and equipment listed on page 37 of the textbook, your teacher will provide you with the following liquids, labelled 1 to 5 (not necessarily in this order):

- starch solution
- vegetable oil
- sodium hydroxide (0.1 mol/L)
- hydrochloric acid (0.1 mol/L)
- sodium thiosulfate (0.1 mol/L)

Note: Your teacher may substitute any of these solutions with another solution.

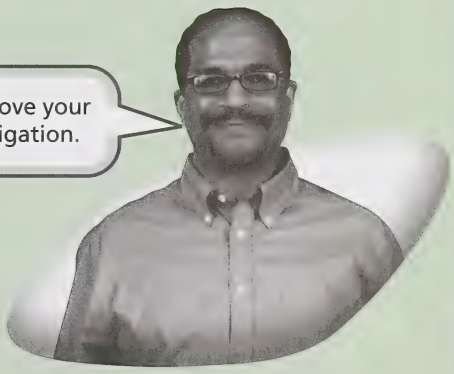
Pay special attention to the safety precautions mentioned. Handle all substances carefully.

Procedure

Follow the instructions as outlined in "Conduct Your Investigation."

32. Complete step 4 of "Conduct Your Investigation." Write your test procedure to all five liquids.

Check your answer with the one on page 57.



Observations

33. Copy and complete the table given on page 37 of the textbook with your observations.

Analysis

34. Which solution was Liquid 1? Liquid 2? Liquid 3? Liquid 4? Liquid 5?

Check your answers with those on page 58.

Part B

35. Complete step 4 of "Conduct Your Investigation."

Check Check your answer with the one on page 58.

The following unlabelled liquids were provided to a student who completed Part A:

- starch solution
- vegetable oil
- hydrochloric acid (0.1 mol/L)
- sodium hydroxide (0.1 mol/L
- sodium thiosulfate (0.1 mol/L)

	Baking Soda	Sodium Chloride	Thymol Blue	Calcium Supplement	Tincture of Iodine	Magnesium Strip (1 cm)
Liquid 1	no response	dissolves	yellow	no response	black	no response
Liquid 2	bubbles appear	dissolves	red	dissolves; liquid bubbles	yellow	liquid bubbles; magnesium corrodes
Liquid 3	dissolves	dissolves	blue	no response	colourless	no response
Liquid 4	no response	no response	sinks	no response	sinks	no response
Liquid 5	no response	dissolves	yellow	does not dissolve	colourless	no response

Looking Back

You have now completed the concepts for this lesson. You analyzed the periodic table and observed how atomic theory is used to describe the structure of the atom. You also observed that elements can combine to form compounds when the atoms of each element gain or lose electrons.

36. Answer questions 9 to 12 of "Check and Reflect" on page 39 of the textbook. **Hint:** For question 11, remember that the atomic number indicates the number of protons the element has.

Check your answers with those on pages 58 and 59.

Go to pages 1 to 3 of Assignment Booklet 1B and answer questions 1 to 11.

Glossary

anion: a negatively charged ion

atomic molar mass: the average molar mass of the atoms of an element including all isotopes

cation: a positively charged ion

compound: a chemical combination of two or more elements in a specific ratio

ductile: able to be drawn into long thin wires

element: a pure substance that cannot be broken down into other substances

energy level: a region near an atom's nucleus that may be empty or may contain electrons

group: a vertical column in the periodic table; also called family

ion: an electrically charged atom or group of atoms

isotope: an atom of the same element containing a different number of neutrons

malleable: able to be beaten or rolled into sheets without crumbling

mass number: an integer equal to the total number of protons and neutrons in the nucleus of the atom

metal: a shiny, malleable, and ductile element

metalloid: an element with properties intermediate between metals and non-metals

non-metal: an element with varying properties that are completely different from metals

There are 17 non-metals in the periodic table.

octet rule: a rule stating that atoms bond in such a way that each valence energy level has eight electrons

period: a horizontal line in the periodic table

valence electron: an electron in the outer energy level of an atom

Suggested Answers

1. Common properties of metals include

   - silver, grey, copper, or gold in colour
   - shiny
   - good conductor of electricity and heat
   - malleable (can be beaten or rolled into sheets)
   - ductile (can be stretched into long wires)
   - solid at room temperature with the exception of mercury

2. Properties that set some metals apart from the others are as follows:

   - Mercury is a liquid at room temperature and normal atmospheric pressure.
   - Sodium is highly reactive in water and air.
   - Platinum and gold are inert.

3. Non-metals have no resemblance to metals.

4. Properties that set some non-metals apart from other non-metals include the following:

   - Eleven are gases, five are solids, and one is a liquid at room temperature and normal atmospheric pressure.
   - There is much variation in colour among non-metals.
   - Some exist in more than one form.
   - Some are highly reactive. For example, fluorine can etch glass. Others, like helium, are unreactive.
   - About half exist in molecule form in normal state (e.g., O_2 and N_2).

5. Metalloids are elements that have properties that are intermediate between metals and non-metals.

6. a. Properties of most of the group 1 elements include the following:

    - They are soft, shiny, and silvery in colour.
    - They are very reactive with water.
    - Compounds tend to be white solids that are soluble in water.

   b. Group 1 elements are called alkali metals.

   c. Although hydrogen is part of column 1, it is a non-metal. The reason why hydrogen is in this group is because it has one electron in its highest energy level. But since hydrogen has only one electron in total, it has properties that are different from other group 1 elements.

> **Note:** You will learn more about energy levels later in this lesson.

7. a. Group 2 elements are called alkaline-earth metals.

   b. Group 2 elements are similar to group 1 elements in that they are shiny and silvery in colour. Also, their compounds tend to be white solids.

8. a. Group 18 elements are called noble gases.
   b. A common property of group 18 elements is that they are very unreactive.

9. a. Group 17 elements are called halogens.
   b. Group 17 elements are very poisonous and react readily with alkali metals to form salts.

10. The atom is the smallest part of an element that still has the properties of that element.

11. electron—negative charge
    proton—positive charge
    neutron—no charge

12. Most of the mass of an atom is located at the centre of the atom in the nucleus.

13. An energy level is a region of space near the nucleus of an atom that may be empty or may contain electrons.

14. Electrons with the greatest amount of energy are found in the energy level **furthest** away from the nucleus. Electrons with the **least** amount of energy are found in the energy level closest to the nucleus.

15. The maximum number of electrons at the first energy level is 2; the maximum number of electrons at the second energy level is 8; and the maximum number of electrons at the third energy level is 8.

16. a. helium—2 protons b. oxygen—8 protons c. iron—26 protons

17. The mass number of an element is equal to the total number of protons and neutrons of the element. **Note:** The periodic table does not contain information about the mass number and, therefore, cannot be used directly to determine the number of neutrons in an atom.

18. There is more than one atomic mass for each element because atoms of the same element may contain different numbers of neutrons.

19. The three isotopes of hydrogen are common hydrogen (with 1 proton and 0 neutrons), deuterium with (1 proton and 1 neutron), and tritium with (1 proton and 2 neutrons).

20. number of neutrons = mass number – atomic number
    = 17 – 8
    = 9

    This isotope of oxygen has 9 neutrons.

21. a. Atomic molar mass is the weighted average mass of the element's naturally occurring isotopes on Earth. The weighted average mass is based on the mass and quantity of each isotope. Atomic molar mass is given under the name for each element in the periodic table.

    b. The atomic molar mass of hydrogen is 1.01 g/mol. It is not exactly 1 because it is the average mass of the three naturally occurring isotopes of hydrogen on Earth.

22. The magnesium ion is a cation, since it has a positive charge.

23. Sodium has a charge of 1+ because it can lose only one electron to become an ion. Magnesium has a charge of 2+ because it can lose two electrons to become an ion.

24. Your drawing should be similar to the following. The element is an isotope of nitrogen.

25. The charge of the ion is 3+. As an element, this ion would have had 13 electrons since it has 13 protons. It then loses 3 electrons, so the ion has a charge of 3+.

26. Metals tend to lose electrons. Non-metals tend to gain electrons.

27. Atoms tend to gain or lose electrons so they have the same number of electrons as the nearest noble gas. This means the outer energy level will have its full compliment of electrons.

28. *Valence number* is the number of electrons an atom of a particular element can gain or lose.

29. a.

Group or Family	Number of Valence Electrons
1	1
2	2
13	3
14	4
15	5
16	6
17	7
18	8

Note: The number of valence electrons is the same as the ones digit in the count number for that group.

b. The number of electrons in the outer energy level is the same for all the elements within each group or family.

c. All the elements in a group or family have similar properties.

d. As you move from left to right in the periodic table, the number of valence electrons increases by 1.

e. The first period has one energy level; the second period has two energy levels; and the third period has three energy levels. Therefore, the period indicates the number of energy levels.

30. Fluorine has seven valence electrons, so it needs to gain one electron to complete its outer energy level. Neon already has eight electrons in its outer energy level, so it does not need to gain or lose any electrons. Magnesium has two valence electrons, so it needs to lose two electrons to have a complete outer energy level with eight electrons.

31. It is difficult for metals to lose more than about three electrons because as electrons are lost in the outer energy level, the remaining electrons are held more tightly to the nucleus.

32. Procedures may vary. A sample procedure is given.

step 1: Place a small quantity of Liquid 1 in six test tubes or in six positions on a spot plate.
step 2: Add each test reagent to a sample of Liquid 1.
step 3: Record your observations in a table similar to the one given on page 37 of the textbook.
step 4: Thoroughly clean the test tubes or spot plate (if needed), and repeat steps 1 to 3 for the remaining unknown liquids.

33. Results may vary. Sample answers are given.

	Baking Soda	Sodium Chloride	Thymol Blue	Calcium Supplement	Tincture of Iodine	Magnesium Strip (1 cm)
Liquid 1	no response	dissolves	yellow	no response	black	no response
Liquid 2	bubbles appear	dissolves	red	dissolves; liquid bubbles	yellow	liquid bubbles; magnesium corrodes
Liquid 3	dissolves	dissolves	blue	no response	colourless	no response
Liquid 4	no response	no response	sinks	no response	sinks	no response
Liquid 5	no response	dissolves	yellow	no response	colourless	no response

34. Liquid 1—starch solution (It turned black with tincture of iodine.)
Liquid 2—hydrochloric acid (The magnesium strip formed bubbles on it.)
Liquid 3—sodium hydroxide (Sodium chloride dissolves in it.)
Liquid 4—vegetable oil (It sinks in water solutions.)
Liquid 5—sodium thiosulfate (It is the last of the unknown liquids.)

35. Refer to the answer to question 32.

36. **Textbook questions 9 to 12 of "Check and Reflect," p. 39**

9. Nitrogen-14 contains seven protons, seven neutrons, and seven electrons. Nitrogen-15 contains seven protons, eight neutrons, and seven electrons. The difference between the two isotopes is that nitrogen-15 has one more neutron.

10. a. Phosphorus will gain three electrons, since its ionic charge is 3–.
b. Sodium will lose one electron, since its ionic charge is 1+.
c. Chlorine will gain one electron, since its ionic charge is 1–.
d. Magnesium will lose two electrons, since its ionic charge is 2+.
e. Iodine will gain one electron, since its ionic charge is 1–.

11.

Element	Mass Number	Number of Protons	Number of Neutrons
calcium	41	20	21
uranium	238	92	146
aluminium	27	13	14
beryllium	9	4	5
neon	19	10	9
iron	53	26	27

12.

Atom or Ion Name	Overall Charge	Number of Protons	Number of Electrons	Symbol	Number of Electrons Lost or Gained
oxygen atom	0	8	8	O	none
oxide ion	2–	8	10	O^{2-}	gained 2
potassium ion	1+	19	18	K^{+}	lost 1
magnesium ion	2+	12	10	Mg^{2+}	lost 2
fluoride ion	1–	9	10	F^{-}	gained 1
calcium ion	2+	20	18	Ca^{2+}	lost 2
aluminium ion	3+	13	10	Al^{3+}	lost 3

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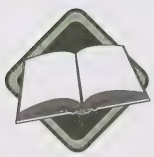
Lesson 2

Naming Ionic and Molecular Compounds

Good communication is important, whether it is just with your friends about the weekend or whether it is with your teacher and/or fellow students about one of the topics you are studying. Effective, understandable communication is required when naming chemicals as well. As a result, the International Union of Pure and Applied Chemistry (IUPAC) was set up to establish rules for naming compounds.

Chalk is a common name for several substances that contain calcium. To distinguish which calcium compound is contained in which substance, scientists have very specific names for each compound. For example, naturally occurring chalk is calcium carbonate. Blackboard chalk is made of gypsum, which is calcium sulphate.



Turn to page 40 of the textbook and read the introductory paragraphs of "Naming Ionic and Molecular Compounds."

1. a. Name two items that contain different forms of chalk.
   b. What is the scientific name for the substance contained in each item?

2. What is IUPAC?

Check your answers with those on page 70.

ionic compound: a compound formed when electrons are transferred from one atom to another

molecular compound: a compound that forms molecules when non-metallic atoms share electrons

There are two classes of compounds: **ionic compounds** and **molecular compounds**. In this lesson you will, first, determine how to name ionic and molecular compounds.

NOTICE

Re: States of Matter

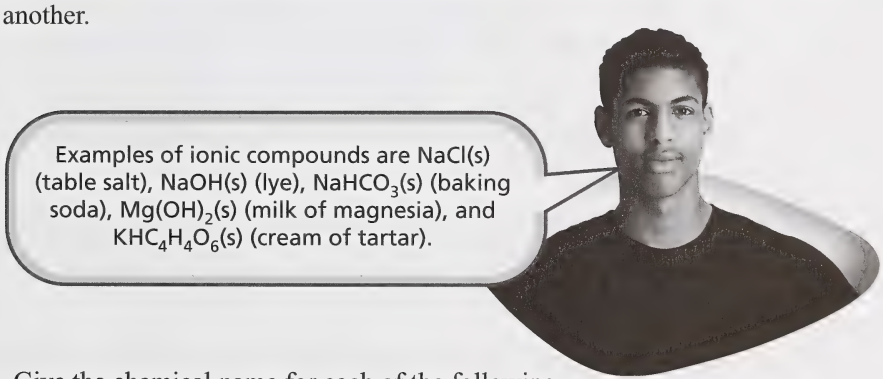
The textbook shows the states of matter of elements and compounds as italicized subscripts. Throughout this course, the IUPAC rule will be followed where states of matter appear in regular type and in brackets.

e.g., Mg(s), H_2(g), and HCl(aq)

When writing chemical formulas, you should follow the IUPAC rule for states of matter.

Ionic Compounds

Ionic compounds are substances formed when electrons transfer from one atom to another.


Examples of ionic compounds are NaCl(s) (table salt), NaOH(s) (lye), $NaHCO_3$(s) (baking soda), $Mg(OH)_2$(s) (milk of magnesia), and $KHC_4H_4O_6$(s) (cream of tartar).

3. Give the chemical name for each of the following common-name substances, and state what each substance is used for. Refer to Table A2.2 at the bottom of page 41 of the textbook.

   a. lye b. baking soda c. milk of magnesia

Check your answers with those on page 70.

crystal lattice: an organized array of ions

formula unit: the smallest amount of an ionic compound that has the composition shown by the chemical formula

Sodium chloride is one of the most common compounds on Earth and one that is quite familiar to everyone. Sodium chloride forms when the sodium atom transfers an electron to the chlorine atom. In sodium chloride, the two sodium and chloride ions combine in an organized array called a **crystal lattice**. There is one sodium ion for each chlorine ion. This combination of one sodium ion and one chlorine ion is called a **formula unit**.

Turn to page 41 of the textbook and study Figures A2.16 and A2.17. These diagrams show how the sodium atom (Na) transfers an electron to the chlorine atom (Cl) to become the stable compound sodium chloride, NaCl(s). The bond formed is called an **ionic bond**. For more information on ionic compounds and ionic bonds, read "Ionic Compounds" on pages 40 and 41 of the textbook.

ionic bond: a bond formed when electrons are transferred between a metal and a non-metal

4. Explain in terms of the octet rule why sodium chloride is a stable compound, whereas sodium atoms and chlorine atoms are not stable?

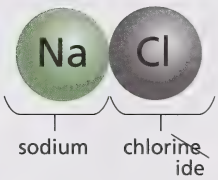
5. Name four properties of ionic compounds.

Check your answers with those on pages 70 and 71.

Naming Ionic Compounds

Ionic compounds are named using the name of the cation and anion that make up the compound. The name of the cation comes first and is used as is; the name of the anion comes second and has the last part changed to *-ide*. For example, in sodium chloride, sodium is the cation element and chlorine is the anion element. The name sodium stays as is and the last part of chlorine is changed to *-ide*, which gives the name sodium chloride. Refer to Figure 2.5 on the right.

Figure 2.5: Sodium Chloride



Turn to page 42 of the textbook and read the two rules for naming ionic compounds at the top of the page. Study the examples of ionic compounds in Table A2.3.

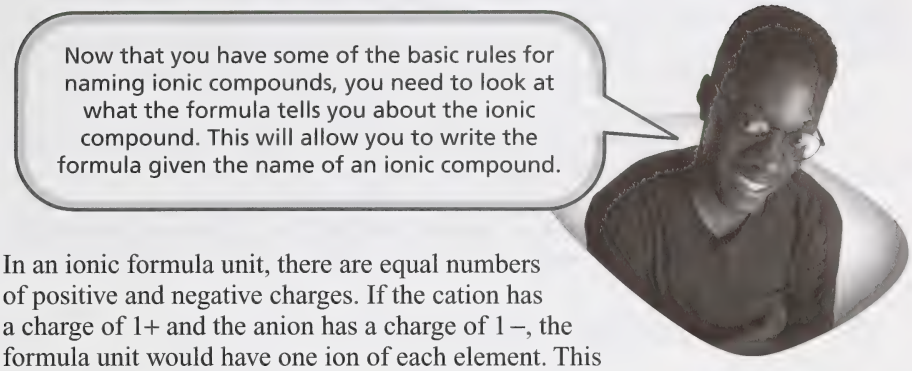
6. Write the symbols for the two ions and the names of the elements that form the compound barium fluoride.

7. Write the ionic name for each of the following ionic compounds.

   a. $K_3N(s)$ b. $AlCl_3(s)$ c. $Li_2O(s)$ d. $CaS(s)$

Check your answers with those on page 71.

Writing Formulas for Ionic Compounds


Now that you have some of the basic rules for naming ionic compounds, you need to look at what the formula tells you about the ionic compound. This will allow you to write the formula given the name of an ionic compound.

In an ionic formula unit, there are equal numbers of positive and negative charges. If the cation has a charge of 1+ and the anion has a charge of 1 –, the formula unit would have one ion of each element. This is what happens when Na^+ and Cl^- combine to form NaCl(s). If the cation has a charge of 3+ and the anion has a charge of 1 –, you need three anion charges to balance the single cation. This occurs when aluminium and chlorine combine to form aluminium chloride. The aluminium ion (Al^{3+}) has a charge of 3+, and the chloride ion (Cl^-) has a charge of 1 –. Therefore, it takes three chloride ions to balance one aluminium ion, making the formula for aluminium chloride $AlCl_3$(s).

For a summary of the steps for writing the formulas of ionic compounds NaCl(s) and $AlCl_3$(s), turn to page 43 of the textbook and study Table A2.4.

8. Write the formula for the following ionic compounds. Use the periodic table on page 30 of the textbook to determine the ionic charge for each ion.

   a. sodium nitride

   b. magnesium fluoride

Check your answers with those on page 71.

In many cases, the number of charges of both the cation and the anion is greater than one. When this happens, you can either determine the total charges needed to balance the formula unit by inspection or you can use the method of lowest common multiple (LCM) to determine the total charges needed.

Read "The Method of Lowest Common Multiple" on page 43 of the textbook. Work through Example Problem A2.1 carefully.

9. Answer question 1.c. of "Practice Problem" on page 43 of the textbook. Explain how you arrived at your answer.

10. Write the formulas for scandium oxide and barium phosphide. Explain how you arrived at your answer.

Check your answers with those on page 71.

multivalent element: an element with more than one stable ion

Some elements have more than one stable ion. For example, iron has Fe^{3+} and Fe^{2+} as stable ions. These elements are called **multivalent elements**. Refer to the periodic table on page 30 of the textbook for other multivalent elements. Many of the multivalent elements occur in periods 4 to 6. If you are given a formula containing a multivalent element, you need to be able to determine which valence number has been used before you can write the name of the compound. It is also necessary to use an upper case Roman numeral to indicate which valence number is used, like iron(III) and lead(IV).

Turn to page 44 of the textbook and read "Compounds with Multivalent Elements." Work through Example Problem A2.2 carefully.

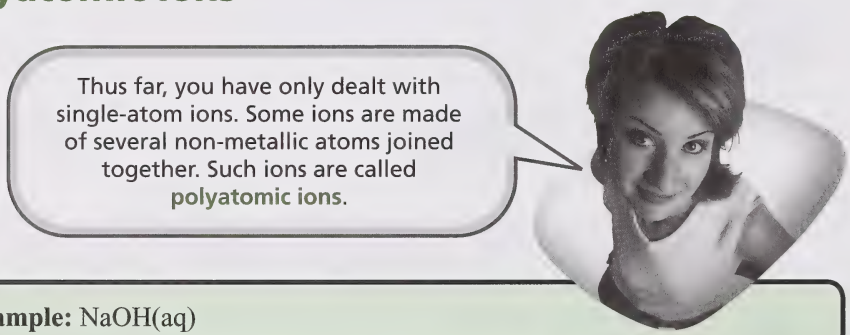
11. How do you know that $FeBr_2$ must be iron(II) bromide and $FeBr_3$ must be iron(III) bromide?

12. In Example Problem A2.2, why wasn't the name of the compound copper(II) nitride?

13. Answer questions 2.b., 2.c., and 2.d. of "Practice Problem" on page 44 of the textbook. Explain how you arrived at your answer.

Check your answers with those on pages 71 and 72.

Polyatomic Ions


Thus far, you have only dealt with single-atom ions. Some ions are made of several non-metallic atoms joined together. Such ions are called polyatomic ions.

polyatomic ion: a charged particle made up of several non-metallic atoms

Example: NaOH(aq)

The cation is Na^+, and the anion is OH^-. The anion is made up of two elements—oxygen and hydrogen—combined to form a single hydroxide ion with a charge of 1−.

Example: $NH_4Cl(aq)$

The cation is NH_4^+, and the anion is Cl^-. The cation is made up of two elements—nitrogen and hydrogen—combined to form a single ammonium ion with a charge of 1+.

The charge for different polyatomic ions cannot be found in the periodic table. Table A2.5 on page 44 of the textbook gives some common polyatomic ions. A special table listing common polyatomic ions can also be found in the Common Polyatomic Ions table on page 491 of the textbook. Refer to this table whenever you need to know the charge of a certain polyatomic ion.

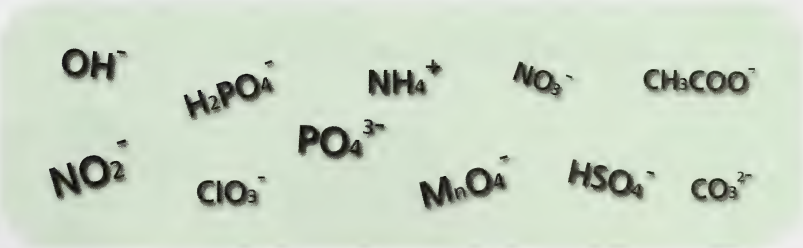
Notice that the name used for the PO_4^{3-} is phosphate. The ending here is *-ate*. The endings *-ate* and *-ite* are the most common suffixes used with polyatomic ions.

Turn to page 45 of the textbook and read the information in "Suffixes for Polyatomic Ions."

14. What is the difference between the suffixes *-ate* and *-ite* when referring to oxygen atoms?

15. What happens when there are more than two similar ions?

Check Check your answers with those on page 72.



Naming compounds containing polyatomic ions is relatively simple. Look at the cation and give the name of the element from which that cation came. Then give the name of the polyatomic ion. You will need to refer to the list of polyatomic ions until you have memorized some of the more common ones.

Study Table A2.7 on page 45 of the textbook to observe some hints for writing names of polyatomic ions.

16. Write the name for the following polyatomic ions.

a. $Cu_3PO_4(s)$ b. $NH_4NO_2(s)$ c. $K_2SO_4(s)$

Check Check your answers with those on page 72.

Writing the formula for a polyatomic ion is similar to writing the formula for other ionic compounds. However, remember to treat the polyatomic ion as one unit with a charge for the group.

Turn to page 46 of the textbook and work through Example Problems A2.3 and A2.4. Follow the steps in each example carefully.

17. How is the charge for the sulfate ion in Example Problem A2.3 obtained?

18. Why are brackets used for the sulfate ion in Example Problem A2.3 but not for the dihydrogenphosphate ion in Example Problem A2.4?

19. Answer questions 3.a., 3.c., 4.a., and 4.b. of "Practice Problems" on page 46 of the textbook. Use the periodic table to determine the charge of the metal ions and Table A2.5 or the Common Polyatomic Ions table to determine the charge of the polyatomic ions. Explain how you determined the formula or compound name for each.

Check Check your answers with those on pages 72 and 73.

Molecular Compounds

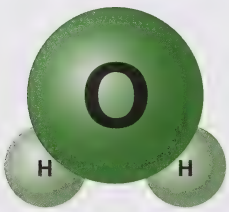


Figure 2.6: Water molecule

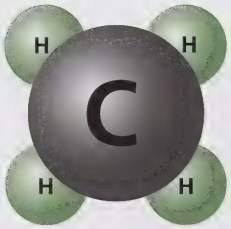


Figure 2.7: Methane molecule

Two compounds of great importance to society are water (H_2O) and methane (CH_4). Neither compound is ionic. Both compounds are made up of molecules and are called molecular compounds. Molecular compounds are made up of independent units made up of fixed numbers of atoms bonded together.

Turn to pages 46 and 47 of the textbook and read "Molecular Compounds."

20. Describe three properties of molecular compounds.

21. What do you notice about the compounds in Table A2.8?

Check Check your answers with those on page 73.

covalent bond: a bond formed when non-metallic atoms share electrons

Unlike ionic compounds where electrons are transferred from one atom to another, atoms in molecular compounds combine by sharing electrons. This sharing of electrons results in a **covalent bond**. The sharing of one pair of electrons forms a single covalent bond; the sharing of two pairs of electrons forms a double covalent bond; and the sharing of three pairs of electrons forms a triple covalent bond.

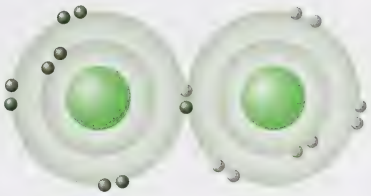


Figure 2.8: Two fluorine atoms share an electron to form a single covalent bond.

For more information on covalent bonds, read "Sharing Electrons—Covalent Bonds" on page 47 of the textbook. Study Figures A2.20, A2.21, and A2.22 closely.

22. How many electrons are shared in a single covalent bond?

23. Name two elements that form covalent bonds between pairs of atoms.

24. What is the name for the covalent bond formed when two nitrogen atoms share electrons?

Check your answers with those on page 73.

Not all elements exist as single atoms. Some exist as two atoms bonded together or three or more atoms bonded together. All metals and noble gases exist as single atoms. Gases, such as nitrogen and oxygen, exist as two atoms bonded together. Elements with three or more atoms include ozone (O_3), phosphorus (P_4), and sulfur (S_8).

For more information on different forms of elements, read "Molecular Elements" at the top of page 48 of the textbook.

25. What is a molecular element?

26. State the names used for the three forms in which elements may exist.

27. What is common about most of the diatomic elements listed in Table A2.9?

28. Read "*re*SEARCH" in the upper left corner of page 48 of the textbook. What is meant by saying that the mercury(I) ion is also a polyatomic ion?

Check your answers with those on page 73.

Insert the Science 10 Multimedia CD into your computer, and view the segment "Electron Dot Diagram." This segment shows how elements combine to form compounds in ionic and molecular bonding. You can stop and start the segment using the buttons provided.

29. a. How does a molecule of sodium form?

    b. How does a crystal of sodium chloride form?

    c. How many electrons transfer when a magnesium ion combines with a sulfide ion?

    d. How do two chlorine atoms combine?

    e. What kind of bond forms when a carbon atom combines with two oxygen atoms?

Check your answers with those on pages 73 and 74.

Naming Molecular Compounds

Molecular compounds that contain two elements are known as binary molecular compounds. Molecular compounds are named in two ways, depending on whether or not the compound contains hydrogen. Binary molecular compounds that do not contain hydrogen are named using IUPAC rules.

Turn to page 48 of the textbook and study Table A2.11. This table provides you with steps to follow when naming binary molecular compounds that do not contain hydrogen. The prefixes used are listed in Table A2.10.

30. How is each atom named in the examples in Table A2.11?

31. What do the prefixes indicate in each example?

32. What is the format for naming a binary molecular compound that does not contain hydrogen?

33. Work through Example Problem A2.5 on page 49 of the textbook. Then complete questions 5.a., 5.c., 5.d., and 5.e. of "Practice Problem."
    Note: To write the formula, given the name of the molecular compound, simply reverse the steps. Explain how you arrived at your answer for each.

Check your answers with those on page 74.

Molecular compounds that contain hydrogen are very common. Some of these compounds have common names that do not follow any rules. These have to be memorized or found in a table. Water, hydrogen peroxide, ammonia, and sucrose are just a few of the molecular compounds containing hydrogen that do not follow specific naming rules.

There is a large class of molecular compounds that contain carbon and hydrogen or carbon and hydrogen and other elements. These are called organic compounds and named according to specific rules. You will cover these compounds in future science courses. Table A2.12, however, gives a few examples of organic compounds. They are methane, propane, methanol and ethanol.

Hydrogen sulfide, $H_2S(g)$, given in Table A2.12, is another exception. When hydrogen reacts with halogens, sulfur, and selenium, the names of the compounds are the same as if they were ionic compounds.

Read "Molecular Compounds That Contain Hydrogen" on page 49 of the textbook. Closely study Table A2.12 for additional examples of these compounds.

34. Name the following molecular compounds given the formula.

    a. $NH_3(g)$ b. $C_{12}H_{22}O_{11}(s)$

35. Write the formula for the following molecular compounds.

    a. propane b. hydrogen peroxide c. ethanol

Check your answers with those on page 74.

Looking Back

You have now completed the concepts for this lesson. You defined ionic and molecular compounds and developed a basic ability in writing formulas and naming these compounds.

36. Answer questions 1, 2, 3.a., 3.d., 3.g., 4.a., 4.d., 5.b., 5.c., 5.e., 6.a., 6.c., 6.d., 6.g., 6.h., 8.a., 8.c., 8.g., 9.c., 9.f., 10.a., 10.g., 10.i., 11.a., and 11.g. of "Check and Reflect" on page 50 of the textbook.

Check your answers with those on page 75.

Go to pages 4 and 5 of Assignment Booklet 1B and answer questions 12 to 19.

Glossary

covalent bond: a bond formed when non-metallic atoms share electrons

crystal lattice: an organized array of ions

formula unit: the smallest amount of an ionic compound that has the composition shown by the chemical formula

ionic bond: a bond formed when electrons are transferred between a metal and a non-metal

ionic compound: a compound formed when electrons are transferred from one atom to another

molecular compound: a compound that forms molecules when non-metallic atoms share electrons

molecular element: an element that forms a molecule made up of only its atoms

multivalent element: an element with more than one stable ion

polyatomic ion: a charged particle made up of several non-metallic atoms

Suggested Answers

1. a. Blackboard chalk and TUMS contain different forms of chalk.

   b. The scientific name for the substance contained in blackboard chalk is calcium sulfate. The scientific name for the substance contained in TUMS is calcium carbonate.

2. **I**nternational **U**nion of **P**ure and **A**pplied **C**hemistry—a body of scientists that establish rules for naming compounds.

3. a. The chemical name for lye is sodium hydroxide. It is used to unclog drains. **Note:** You may be more familiar with it as the trade name Draino or Liquid Plumber.

   b. The chemical name for baking soda is sodium hydrogencarbonate. It is used in baking where it reacts with an acid like cream of tartar to give off CO_2 and make baked goods rise.

   c. The chemical name for milk of magnesia is magnesium hydroxide. It is used as an antacid and laxative. (Note the difference in properties between sodium hydroxide and magnesium hydroxide—one is used to clean drains and is quite poisonous, and the other is used to settle an upset stomach.)

4. Sodium chloride is stable because the outer energy level on each ion is full. When sodium transfers an electron to chlorine, it ends up with 8 electrons in its outer energy level. When chlorine receives an electron, it ends up with 8 electrons in its outer energy level. Eight electrons in the outer energy level of an atom is a full energy level.

5. Four properties of ionic compounds are as follows:

   - solid at room temperature
   - some can be heated to a very high temperature without decomposing or burning
   - tend to dissolve in water
   - solutions of ionic compounds conduct electricity

6. The symbol for the cation is Ba^{2+}, and the name of the element from which this ion comes is barium. The symbol for the anion is F^-, and the name of the element from which this ion comes is fluorine.

7. a. potassium nitride b. aluminium chloride c. lithium oxide d. calcium sulfide

8. a. $Na_3N(s)$ b. $MgF_2(s)$

9. **Textbook questions 1.c. of "Practice Problem," p. 43**

   1. c. $Ca_3P_2(s)$

      The calcium ion (Ca^{2+}) has a charge of 2+, and the phosphorus ion (P^{3-}) has a charge of 3–. The lowest common multiple of 2 and 3 is 6. Dividing 6 by 2 gives 3 calcium ions, and dividing 6 by 3 gives 2 phosphide ions.

10. $Sc_2O_3(s)$

    The scandium ion (Sc^{3+}) has a charge of 3+, and the oxygen ion (O^{2-}) has a charge of 2–. So, the lowest common multiple is 6. Dividing 6 by 3 gives 2 ions of scandium, and dividing 6 by 2 gives 3 ions of oxygen.

    $Ba_3P_2(s)$

    The barium ion (Ba^{2+}) has a charge of 2+, and the phosphorus ion (P^{3-}) has a charge of 3–. So, the lowest common multiple is 6. Dividing 6 by 2 gives 3 ions of barium, and dividing 6 by 3 gives 2 ions of phosphorus.

11. From the periodic table you know that an iron (Fe) ion could have a 2+ or 3+ charge. You also know that the bromide ion (Br^-) has a 1– charge. If there are two bromide ions in the formula, the charge on the iron ion must be 2+ to balance the 2– charge of the two bromide ions. The name given to the compound is iron(II) bromide. If there are three bromide ions in the formula, the charge on the iron ion must be 3+ to balance the 3– charge of the three bromide ions. The name given to the compound is iron(III) bromide.

12. The name of the compound in Example Problem A2.2. is not copper(II) nitride because the ion for copper(II) would be the one with the 2+ charge for each ion. Three Cu^{2+} would give a total charge of 6+ for the cations. The anion would then have to have a total charge of 6–, but the total charges for the nitride ion is only 3–. Three Cu^+ gives a total charge of 3+ to balance the 3– charge of the nitride ion.

13. **Textbook questions 2.b., 2.c., and 2.d. of "Practice Problem," p. 44**

    2. b. lead(IV) oxide

       The charge of an oxygen ion is 2−. There are two oxygen ions, so the total charge is 4−. To balance this charge, the total charge on the lead ion needs to be 4+. Because there is only one lead (Pb) ion in the formula unit, the charge on that ion must be 4+. Therefore, lead(IV) is used.

       c. nickel(III) sulfide

       The charge of a sulfide ion is 2−. There are three sulfide ions, so the total charge is 6−. To balance this charge, the total charge on the nickel ions need to be 6+. Because there are 2 nickel ions, the charge on each nickel ion must be 3+. Therefore, nickel(III) is used.

       d. copper(II) fluoride

       The charge of a fluoride ion is 1−. There are two fluoride ions, so the total charge is 2−. To balance this charge, the total charge on the copper ion needs to be 2+. Because there is only one copper ion, the charge on this ion must be 2+. Therefore, copper(II) is used.

14. The suffix *-ate* means more oxygen atoms are part of the ion. The suffix *-ite* mean s fewer oxygen atoms are part of the ion.

15. When there are more than two similar ions, the prefixes *per-* and *hypo-* are used. The prefix *per-* is used when there is one more oxygen atom than in the *-ate* ion. The prefix *hypo-* is used when there is one less oxygen atom than in the *-ite* ion.

16. a. copper(I) phosphate b. ammonium nitrite c. potassium sulfate

17. The charge of the sulfate ion is obtained by referring to Table A2.5 on page 44 of the textbook or by referring to the Common Polyatomic Ions table on page 491.

18. Brackets are used around the sulfate ion in Example Problem A2.3 because there are three sulfate ions in the formula. Brackets are not required around the dihydrogenphosphate ion because there is only one dihydrogenphosphate ion in the formula.

19. **Textbook questions 3.a., 3.c., 4.a., and 4.b. of "Practice Problems," p. 46**

    3. a. $Ba(OH)_2(s)$

       The barium ion (Ba^{2+}) has a charge of 2+. A hydroxide ion (OH^-) has a charge of 1−. To balance the charges, two hydroxide ions are needed for each barium ion.

c. $CuMnO_{4(s)}$

The charge of the permanganate ion (MnO_4^-) is 1 – . The charge of the copper(I) ion(Cu^+) is 1 +. The charges are balanced with one copper(I) ion and one permanganate ion.

4. a. gold(III) nitrate

The charge of a nitrate ion (NO_3^-) is 1 –. There are three nitrate ions, so a charge of 3 + is needed on the gold ion. The name gold(III) nitrate is used to show that Au^{3+} is the cation.

b. ammonium phosphate

NH_4^+ is the ammonium ion, and PO_4^{3-} is the phosphate ion.

20. Three properties of molecular compounds are as follows:

- can be solid, liquid, or gas at room temperature and normal atmospheric pressure
- tend to be poor conductors of electricity
- many do not dissolve easily in water

21. The molecular compounds in the table all contain carbon and hydrogen. Three out of the four also contain oxygen.

22. Two electrons are shared in a single covalent bond—one from each atom.

23. Chlorine and nitrogen are two elements that share electrons to form covalent bonds.

24. A triple covalent bond is formed when two nitrogen atoms share electrons.

25. A molecular element is an element that forms a molecule made up of one type of atom.

26. The names used for the three forms elements may exist are monatomic (single atom), diatomic (two atoms combined), and polyatomic (more than two atoms combined).

27. Most of the diatomic elements listed are gases at standard temperature and normal atmospheric pressure. The exceptions are bromine and iodine.

28. Mercury exists with two atoms combined in the mercury(I) ion. The mercury(I) ion is written as Hg_2^+ while the mercury(II) ion is written as Hg^{2+}.

29. a. A molecule of sodium chloride forms when the sodium ion transfers an electron to the chloride ion.

b. Many sodium ions each transfer an electron to a chloride ion, forming sodium chloride molecules. These sodium chloride molecules, together, form a crystal of sodium chloride.

c. Two electrons transfer when a magnesium ion combines with a sulfur ion.

d. Two chlorine atoms combine by sharing electrons to form a molecule of chlorine.

e. Two double covalent bonds form when a carbon atom combines with two oxygen atoms.

30. The first element is given its name from the periodic table. The second element is named starting with its name from the periodic table and changing the ending to *-ide*.

31. The prefix indicates how many atoms of the element are present in the formula unit.

32. The format is as follows:

    prefix + name of first element followed by prefix + name of second element with ending changed to *-ide*

33. **Textbook questions 5.a., 5.c., 5.d., and 5.e. of "Practice Problem," p. 49**

    5. a. carbon dioxide

       The name of the C atom is carbon. This is the first part of the name. Because there is only one atom, no prefix is needed. The name of the O atom is oxygen, which becomes oxide. This is the second part of the name. Because there are two oxygen atoms, the prefix *di-* is added to the front of this part of the name.

    c. phosphorous trichloride

       The name of the P atom is phosphorus. This is the first part of the name. Because there is only one atom, no prefix is needed. The name of the Cl atom is chlorine, which becomes chloride. This is the second part of the name. Because there are three chlorine atoms, the prefix *tri-* is added to the front of this part of the name.

    d. $OF_2(g)$

       The symbol for oxygen is O, which is the first part of the formula. Because there is no prefix, there is no subscript. The symbol for fluoride is F, which is the second part of the formula. The prefix *di-* indicates that there are two atoms of fluorine in the formula unit; so the subscript 2 is needed.

    e. $N_2S_4(s)$

       The symbol for nitrogen is N, which is the first part of the formula. The prefix *di-* indicates there are two nitrogen atoms; so the subscript 2 is needed. The symbol for sulfide is S, which is the second part of the formula. The prefix *tetra-* indicates that there are four sulfur atoms; so the subscript 4 is needed.

34. a. ammonia b. sucrose

35. a. $C_3H_8(g)$ b. $H_2O_2(l)$ c. $C_2H_5OH(l)$

36. **Textbook questions 1, 2, 3.a., 3.d., 3.g., 4.a., 4.d., 5.b., 5.c., 5.e., 6.a., 6.c., 6.d., 6.g., 6.h., 8.a., 8.c., 8.g., 9.c., 9.f., 10.a., 10.g., 10.i., 11.a., and 11.g. of "Check and Reflect," p. 50**

1. a. An *ion* is an electrically charged atom or group of atoms. Examples include the sodium ion (Na^+), ammonium ion (NH_4^+), and carbonate ion (CO_3^{2-}).

   b. A *cation* is a positively charged ion. Examples include the sodium ion (Na^+), ammonium ion (NH_4^+), and hydrogen ion (H^+).

   c. An *anion* is a negatively charged ion. Examples include the oxide ion (O^{2-}), nitrate ion (NO_3^-), and hydroxide ion (OH^-).

   d. A *polyatomic ion* is a charged particle with more than one non-metallic atom. Examples include the ammonium ion (NH_4^+) and carbonate ion (CO_3^{2-}).

   e. A *multivalent metal* is a metal with more than one stable ion. Examples included iron (Fe^{2+} and Fe^{3+}) and titanium (Ti^{3+} and Ti^{4+}).

2. 1–mono 2–di 3–tri 4–tetra 5–penta
   6–hexa 7–hepta 8–octa 9–ennea (nona) 10–deca

3. a. Na^+ d. Cu^{2+} g. ClO_3^-

4. a. aluminium ion d. nickel(III) ion

5. b. ammonia c. water e. hydrogen fluoride

6. a. ionic c. molecular d. ionic g. ionic h. molecular

8. a. aluminium chloride c. sodium nitride g. lead(IV) nitrate

9. c. $Mg(SCN)_2(s)$

   (Refer to the Common Polyatomic Ions table on page 491 of the textbook.)

   f. $CrCl_3(s)$

10. a. $N_2O_4(g)$ g. $SO_3(g)$ i. $NH_3(g)$

11. a. carbon tetrabromide g. dinitrogen trioxide

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Lesson 3

Properties and Classification of Ionic and Molecular Compounds

Everybody is familiar with sodium chloride—commonly known as table salt. What are some of the things that help you identify sodium chloride? Well, it's a solid, white in colour, tastes salty, and its particles appear to have sharp edges. These are some common properties that help people to identify table salt.

Scientists note other properties to help identify salt or any other compound. Both ionic and molecular compounds have specific properties that can be used to identify them.

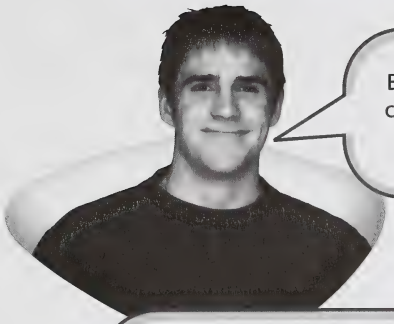
Turn to page 51 of the textbook and read the introductory paragraphs of "Properties and Classification of Ionic and Molecular Compounds." Also, read "*info*BIT" in the upper right corner of the page regarding DNA.

1. The reading in the textbook mentioned the compounds sodium chloride, magnesium sulfate, and lead(II) iodide. How do these compounds differ? How are they similar?

2. How do molecular compounds differ from ionic compounds?

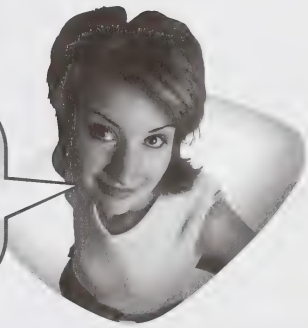
3. Is DNA an ionic or molecular compound?

Check your answers with those on page 85.



Based on its properties, how can you tell if a compound is ionic or molecular?

The reading on page 51 of the textbook stated that ionic compounds in a solution conduct electricity, whereas molecular compounds in a solution do not. There are a number of tests that can be done to determine whether a substance is ionic or molecular.



Inquiry Lab

Ionic or Molecular?

Turn to pages 52 and 53 of the textbook and read the entire activity.

4. Write a hypothesis that answers "The Question." Refer to "Skill Practice: Writing a Hypothesis" on page 51 of the textbook if you need help.

Check your answer with the one on page 85.

If you have access to a supervised laboratory, do **Part A**. If you do not have access to a supervised laboratory do **Part B**.

Part A

Once you have obtained the materials and equipment required for this activity, follow the procedure as outlined on pages 52 and 53 of the textbook.

Pay close attention to the safety precautions mentioned. Both potassium iodide and naphthalene flakes are toxic. Naphthalene is also flammable.

5. Copy and complete a table similar to the one given at the bottom of page 52 of the textbook.

6. a. Answer questions 1 to 6 of "Analyzing and Interpreting" on page 53 of the textbook.

   b. Use the template at the back of this Student Module Booklet to complete a lab report for this activity.

To review what you have learned about writing a lab report in previous science courses, read "Student Reference 11: Writing Reports" on pages 486 and 487 of the textbook. This reference describes all the steps you should follow in writing a lab report.

Check your answers with those on pages 86 and 87.

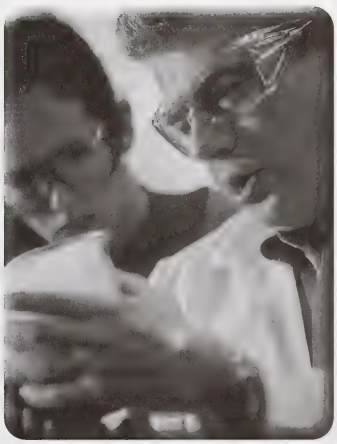
Part B

Use the following observations to answer question 6 in Part A.

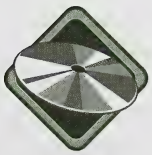
Substance	Appearance	Hardness	Odour	Solubility in Water	Conductivity	Melting Point
sodium chloride	crystalline	hard	no odour	soluble	conducts	does not melt
sugar	crystalline	hard	no odour	soluble	does not conduct	melts quickly (#3)
potassium iodide	crystalline	hard	no odour	soluble	conducts	does not melt
beeswax	waxy	soft	no odour	not soluble	does not conduct	melts quickly (#1)
Epsom salts	crystalline	hard	no odour	soluble	conducts	does not melt
naphthalene flakes	waxy	soft	has an odour	not soluble	does not conduct	melts quickly (#2)

To sum up the preceding Inquiry Lab, you found that some compounds have high melting points, are soluble in water, and their solutions conduct electricity. These compounds are ionic and include Epsom salts, sodium chloride, and potassium iodide. You also found that other compounds have low melting points, are not soluble in water, and their solutions do not conduct electricity. These compounds are molecular and include bees wax and naphthalene flakes. The ionic compounds you analyzed have a crystalline appearance and the molecular compounds appear waxy. However, sugar has a crystalline appearance and is soluble in water; but it has a low melting point. The sugar solution also did not conduct electricity. Sugar shares properties of both ionic compounds and molecular compounds, so sugar was placed between. That being said, sugar is actually classified as a molecular compound.

Properties of Ionic Compounds



You will now look at how to identify an ionic compound from its formula and analyze the properties of ionic compounds in more detail.



Insert the Science 10 Multimedia CD into your computer. Look at the picture of sodium burning in chlorine gas in the Photo Gallery for Module 1. This picture shows an example of the formation of an ionic compound—sodium chloride, NaCl(s).

Now, turn to pages 53 to 55 of the textbook and read "Identifying Ionic Compounds" and "Properties of Ionic Compounds." Examine Table A2.13 and study Figures A2.23 and A2.25 closely.

7. How can you identify an ionic compound from its formula?

8. How consistent are the properties for the four ionic compounds given in Table A2.13?

9. Why do ionic compounds tend to have high melting points?

10. What do the crystals of an ionic compound look like?

11. In terms of positive and negative charges, explain how breaking a crystal structure results in a flat surface. Refer to Figure A2.23.

12. Why do ionic compounds dissolve in water?

Check your answers with those on page 88.

For more information on how ionic compounds dissolve in water, go to the Addison Wesley Science 10 website.

http://www.scienceman.com/science10

Once there, click on "Unit A: Tech Ideas," and scroll down to Field Test Text Pages 51–52. Click on the link *The Solubility of NaCl* to view an interactive display of how NaCl dissolves in water.

13. Use the interactive display to explain how NaCl dissolves in water.

14. Why are ionic compounds good conductors of electricity?

15. Why is a solution of 2 g of NaCl dissolved in water twice as conductive as a solution with 1 g of NaCl dissolved in water?

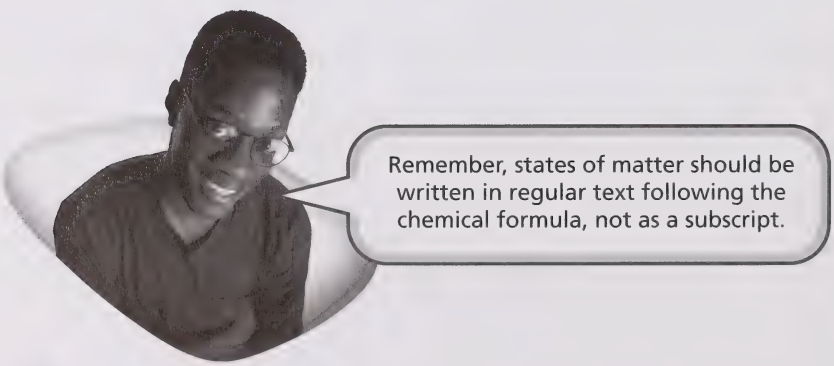
Check your answers with those on page 88.

To observe the conductivity of a number of ionic compounds dissolved in water, go back to the Addison Wesley Science 10 website (as shown previously). Click on "Unit A: Tech Ideas," and scroll down to Field Test Text Page 55. Click on the link *Ionic Solutions Conduct Electricity* to view an interactive display showing ionic solutions conducting electricity.

solubility: a measure of how well a solute dissolves in a particular solvent

The properties of **solubility** and conductivity are of great importance to chemists. Chemists often check the solubility or conductivity of a compound when analyzing it. Scientists test compounds for solubility many times and then put the results into a table of solubility. A table of solubility shows how soluble common ionic compounds are in water. The solubility is rated as very soluble or slightly soluble. In addition, the solubility of the ions is given as to whether it is with all or most compounds containing the ion or if it is only with specific ions.

Turn to pages 57 and 58 of the textbook and read "Solubility of Ionic Compounds." Study Table A2.14 closely.



16. a. How is the chemical formula written for a compound that is very soluble in water?

    b. How is the chemical formula written for a compound that is only slightly soluble in water?

17. Table A2.14 states that all compounds that contain group 1 elements are very soluble in water. Which type of elements are in group 1?

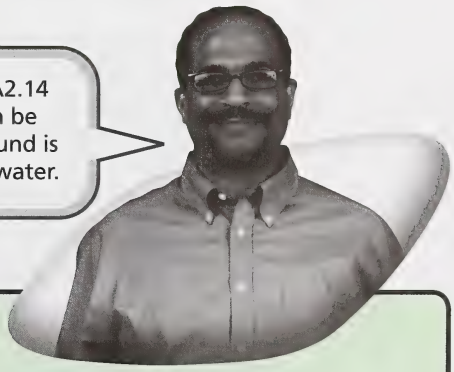
18. Which ions not in group 1 are contained in compounds that are all very soluble in water?

19. Why is it fortunate that phosphates and carbonates are only slightly soluble in water?

20. What is a precipitate?

Check your answers with those on page 89.

A solubility table, such as Table A2.14 on page 57 of the textbook, can be used to predict whether a compound is very soluble or slightly soluble in water.



Example 2.1

Is aluminium chloride, $AlCl_3$, soluble in water?

Solution

step 1: Locate the Cl^- ion in Table A2.14.

The Cl^- ion is located in the fifth column of the table.

step 2: Check the solubility of compounds with Cl^-.

According to the table, most compounds with the Cl^- ion are very soluble in water. The aluminium ion (Al^{3+}) is not listed as one of the elements with which compounds with Cl^- ion are only slightly soluble.

Therefore, aluminium chloride is very soluble in water.

Example 2.2

Determine the solubility of silver nitrate, $AgNO_3$. In the formula, write *(aq)* if the compound is very soluble in water or *(s)* if it is only slightly soluble in water.

Solution

step 1: Locate the nitrate ion (NO_3^-) in Table A2.14.

step 2: Check the solubility of compounds with NO_3^-.

> Compounds containing the nitrate ion (NO_3^-) are all very soluble in water.

Therefore, silver nitrate is very soluble in water and should be written as $AgNO_3(aq)$.

Example 2.3

Determine the solubility of lead(II) bromide, $PbBr_2$. In the formula, write *(aq)* if the compound is very soluble in water or *(s)* if it is only slightly soluble in water.

Solution

According to Table A2.14, most compounds containing bromide ions (Br^-) are very soluble in water. However, bromide compounds with lead(II) ions (Pb^{2+}) are only slightly soluble in water.

Therefore, lead(II) bromide is only slightly soluble in water and should be written as $PbBr_2(s)$.

21. Answer questions 1.a., 1.c., 1.e., 1.h., 1.j., 2.a., 2.d., and 2.g. of "Minds On . . . Using the Solubility Chart" on page 58 of the textbook.

Check your answers with those on pages 89 and 90.

Properties of Molecular Compounds

Do you ever wonder why water freezes at 0°C, yet gasoline remains a liquid even on the coldest winter day?

Molecular compounds exhibit special properties due to the type of bonding that exists between individual atoms and between molecules of the compound. In molecular compounds, often the bonds between the individual atoms are very strong; but the bonds between molecules are weak. A property, such as a low melting point, is a direct result of this weak bond between molecules.

Turn to pages 58 and 59 of the textbook and read "Properties of Molecular Compounds." Examine Table A2.15 closely.

22. How is a water molecule different from a hydrogen peroxide molecule?

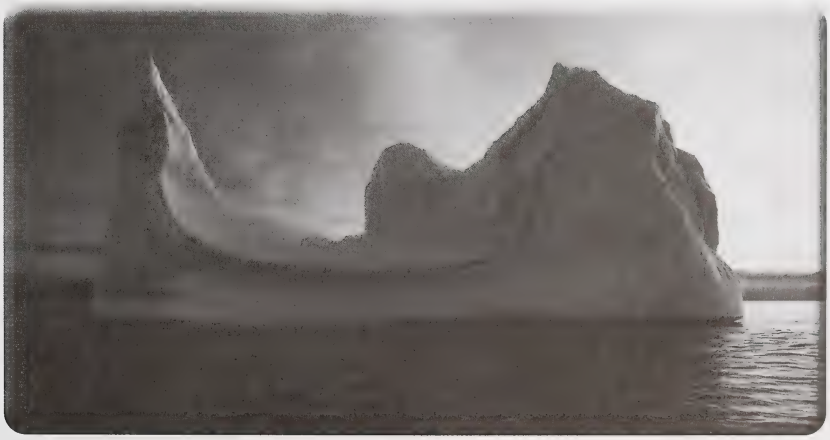
23. Which properties of molecular compounds in Table A2.15 help most when distinguishing molecular compounds from ionic compounds?

24. Name one molecular compound that forms crystals.

25. Read "*info*BIT" on page 59 of the textbook. Why is methane, $CH_4(g)$, sometimes called swamp gas?

Check your answers with those on page 90.

Why does ice float? In most cases as a substance changes from gas to liquid to solid, the substance becomes denser. If the solid state is the densest, it should sink; yet, ice floats. This is just one of the special properties of water that makes it a unique substance.



To learn more about the special properties of water, turn to pages 60 and 61 of the textbook and read "Special Properties of Water."

26. What is the most important special property of water?

27. Use Figure 2.9 to explain how liquid water molecules are held together.

28. How does the polarity of water molecules affect the boiling point of water? Compare the boiling point of water with those of other compounds containing hydrogen in Table A2.16.

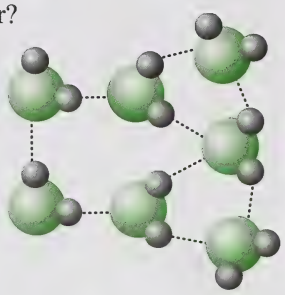


Figure 2.9: Hydrogen bonds in liquid water molecules

29. How does the polarity of water molecules relate to living systems on Earth?

30. Why do many snowflakes have six sides?

31. How does the structure of ice compare to the structure of liquid water?

32. a. Why does ice float?
    b. Why is the fact that ice floats important to fish in a lake?

Check Check your answers with those on pages 90 and 91.

Looking Back

You have now completed all the concepts for this lesson. You identified ionic and molecular compounds and described some of their properties. You also described the special properties of water and identified some of the conditions that result from these special properties.

33. Answer questions 1, 3, 5, 6.a., 6.d., 8.b., and 8.e. of "Check and Reflect" on page 61 of the textbook.

Check Check your answer with the one on page 91.

Go to ... Go to pages 5 and 6 of Assignment Booklet 1B and answer questions 20 to 27.

Glossary

electrolyte: a solution that conducts electricity

polar molecule: a molecule that has a positive charge at one end and a negative charge at the other end

precipitate: a solid of low solubility that forms from mixing solutions

solubility: a measure of how well a solute dissolves in a particular solvent

Suggested Answers

1. The compounds are different in that sodium chloride is edible, magnesium sulfate is used in fertilizer or for soaking feet, and lead(II) iodide is poisonous.

   The compounds are similar in the following ways:

   - They all melt at a temperature above 800°C.
   - They all form crystals with flat surfaces and well-defined edges.
   - None of the solid compounds conduct electricity, but their solutions conduct electricity.
   - All three compounds are ionic. That is, they are made up of positive and negative ions.

2. Molecular compounds differ from ionic compounds in the following ways:

   - Molecular compounds melt at lower temperatures (many melt at temperatures below 250°C).
   - Molecular compounds do not contain metals.
   - Molecular compounds do not conduct electricity in solid form nor in solution form.

3. DNA is a molecular compound.

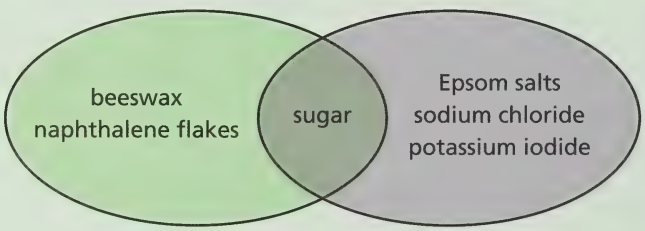
4. If a substance has a certain set of properties, it is ionic. If a substance has another set of properties, it is molecular.

5. Your completed table should be similar to the following.

Substance	Appearance	Hardness	Odour	Solubility in Water	Conductivity	Melting Point
sodium chloride	crystalline	hard	no odour	soluble	conducts	does not melt
sugar	crystalline	hard	no odour	soluble	does not conduct	melts quickly (#3)
potassium iodide	crystalline	hard	no odour	soluble	conducts	does not melt
beeswax	waxy	soft	no odour	not soluble	does not conduct	melts quickly (#1)
Epsom salts	crystalline	hard	no odour	soluble	conducts	does not melt
naphthalene flakes	waxy	soft	has an odour	not soluble	does not conduct	melts quickly (#2)

6. a. **Textbook questions 1 to 6 of "Analyzing and Interpreting," p. 53**

1. to 5. Your Venn diagram should be similar to the following. Note that sugar is placed in the overlapping part since it has some properties similar to beeswax and some similar to Epsom salts.



6. The compounds listed with the bees wax are molecular, and those listed with the Epsom salts are ionic.

b. **Title:** Ionic or Molecular?

Hypothesis: If a substance has a certain set of properties, it is ionic. If a substance has another set of properties, it is molecular.

Materials and Equipment:

- sodium chloride
- beeswax
- 2 hot plates
- conductivity meter
- sugar
- Epsom salts
- 6 test tubes
- potassium iodide
- naphthalene flakes
- aluminium foil

Procedure: Your procedure should be similar to the one given on page 52 of the textbook.

Observations:

Substance	Appearance	Hardness	Odour	Solubility in Water	Conductivity	Melting Point
sodium chloride	crystalline	hard	no odour	soluble	conducts	does not melt
sugar	crystalline	hard	no odour	soluble	does not conduct	melts quickly (#3)
potassium iodide	crystalline	hard	no odour	soluble	conducts	does not melt
beeswax	waxy	soft	no odour	not soluble	does not conduct	melts quickly (#1)
Epsom salts	crystalline	hard	no odour	soluble	conducts	does not melt
naphthalene flakes	waxy	soft	has an odour	not soluble	does not conduct	melts quickly (#2)

Conclusion: The sodium chloride, potassium chloride, and Epsom salts have similar properties. These substances are classified as ionic.

The beeswax and naphthalene flakes have similar properties. These substances are classified as molecular.

The sugar has some properties of both types of compounds.

7. The formula of an ionic compound always has a positive metal ion or positive ammonium ion as the first part and a negative non-metal ion or negative polyatomic ion as the second part.

8. For each property given, at least three of the four compounds are consistent.

    - white in colour (except $PbI_2(s)$)
    - solid at room temperature
    - high melting point (in particular those with metal ions)
    - not malleable
    - soluble in water ($PbI_2(s)$ is slightly soluble.)
    - non-conductor as a solid
    - conductor in solution ($PbI_2(s)$ is a weak conductor.)

9. Ionic compounds tend to have high melting points because strong, attractive forces between the cations and anions hold the ions in an organized lattice of positive and negative ions.

10. The crystals of an ionic compound have flat surfaces. When larger chunks of a solid ionic compound are broken up, the smaller chunks look much like the larger chunks.

11. Normally, the positive and negative charges alternate in a crystal structure. When the crystal is hit, the positive charges momentarily line up beside other positive charges, causing a weakness in the bonding along that line and resulting in the crystal splitting apart along the line.

12. Water molecules are polar. That is, water molecules have a positive end and a negative end. Ionic compounds are made of cations and anions. These charged ions are strongly attracted to water molecules. This attraction causes the ions to break away from the crystal and go into solution.

13. Initially, the ions in solid sodium chloride are packed tightly together. The solid sodium chloride is placed in water, where the water molecules are in rapid motion. The ions in the solid sodium chloride are also in motion but to a lesser extent. The sodium and chloride ions along the surface of the solid move out into the water slightly but are attracted back to the sodium chloride. Some of the faster moving water molecules collide with the sodium or chloride ions at the surface of the solid sodium chloride and cause the ions to move into the water molecules. The attractions between the positive or negative end of the water molecule and the oppositely charged ion cause these ions to stay within the water. Eventually, all of the sodium and chloride ions are knocked away from the surface of the solid sodium chloride and into the water.

14. Ionic compounds are good conductors of electricity because they form positive and negative ions in a solution.

15. There are twice as many positive and negative ions in the solution with 2 g of NaCl, making it twice as conductive.

16. a. The chemical formula for a compound that is very soluble in water has *(aq)* written at the end.

    b. The chemical formula for a compound that is only slightly soluble in water has *(s)* written at the end.

17. The alkali metals are in group 1. These elements are very reactive and include lithium, sodium, potassium, rubidium, cesium, and francium.

18. Ions not in group 1 that are often in compounds that are very soluble in water are NH_4^+, $H_3O^+(H^+)$, ClO_3^-, NO_3^-, ClO_4^-, Cl^-, Br^-, and I^-.

19. Phosphates and carbonates are present in eggshells, clamshells, bones, and teeth. All of these things are usually in contact with water. If phosphates and carbonates were very soluble in water, these things would dissolve.

20. A precipitate is a solid with low solubility that sometimes forms when two ionic solutions are mixed.

21. **Textbook questions 1.a., 1.c., 1.e., 1.h., 1.j., 2.a., 2.d., and 2.g. of "Minds On . . . Using the Solubility Chart," p. 58**

    1. a. $(NH_4)_2S(aq)$

       All compounds with ammonium ions (NH_4^+) are very soluble in water.

       c. $PbSO_4(s)$

       Most compounds with sulfate ions (SO_4^{2-}) are very soluble in water. However, sulfate compounds with lead(II) ions (Pb^{2+}) are only slightly soluble in water.

       e. $Fe(OH)_3(s)$

       Hydroxide (OH^-) compounds with iron(III) ions (Fe^{3+}) are only slightly soluble in water.

       h. $Na_3PO_4(aq)$

       Phosphate (PO_4^{3-}) compounds with group 1 metals, which includes sodium, are very soluble in water.

       j. $AgCH_3COO(s)$

       Acetate (CH^3COO^-) compounds with silver ions (Ag^+) are only slightly soluble in water.

2. a. Potassium carbonate is very soluble in water. Carbonate (CO_3^{2-}) compounds with group 1 metals, which includes potassium, are very soluble in water.

   d. Barium hydroxide is very soluble in water. Hydroxide (OH^-) compounds with barium ions (Ba^{2+}) are very soluble in water.

   g. Lead(IV) bromide is very soluble in water. Bromide (Br^-) compounds with lead(IV) ions (Pb^{4+}) are very soluble in water.

22. A water molecule has one oxygen atom; a hydrogen peroxide molecule has two oxygen atoms.

23. The properties low melting point (under 250°C) and non-conductor when the compound is a solid and when in solution help most when distinguishing molecular compounds from ionic compounds.

24. Snowflakes (ice crystals) and sugar are molecular compounds in crystal form.

25. Methane is sometimes called swamp gas because it is produced in swamps by rotting vegetation. It is also the fossil fuel widely used in Alberta for industry and heating homes.

26. The most important special property of water is that it is polar. That is, each water molecule has a slightly positive and a slightly negative end.

27. The slightly positive end of one molecule is attracted to the slightly negative end of another molecule. This attraction is what holds liquid water molecules together.

28. Table A2.16 shows other compounds similar to water—each compound has two hydrogen atoms attached to another atom. The non-polar molecules (H_2Se and H_2S) have low boiling points compared to water.

29. The polarity of water molecules is responsible for the high boiling point of water, which results in water existing as a liquid on Earth. Liquid water is essential to all living systems on Earth. Liquid water is also important to living systems in that oceans and large lakes regulate the temperature of the Earth. The polarity of water molecules is also responsible for the density of ice being less than the density of water, so ice will float on water. This prevents oceans, lakes, and rivers from freezing solid, allowing all aquatic life to survive in cold climates.

30. Many snowflakes have six sides because when liquid water turns to ice, the molecules spread out and line up in a three-dimensional array that contains six-sided rings. Six-sided snowflakes tend to form between temperatures of about -10°C to -15°C.

31. In ice, the molecules are not free to move. They are fixed in certain shapes. In water, the molecules are free to move.

32. a. Ice floats because it is less dense than liquid water. The ordered structure of molecules in ice takes more space. So, there are fewer molecules in ice than in water for a given volume. Fewer molecules per unit volume results in a lower density.

    b. It is important for fish that ice floats because water remains a liquid beneath the ice where fish can survive the winter. If ice did not float, the water in a lake would freeze from the bottom up and the lake would become an entire mass of ice, killing all the fish.

33. **Textbook questions 1, 3, 5, 6.a., 6.d., 8.b., and 8.e. of "Check and Reflect," p. 61**

    1. The formula of an ionic compound begins with a metal ion or an ammonium ion.

    3. The formula of a molecular compound does not contain any metals.

    5. Water being a polar molecule means that each water molecule has a slightly positive charge at one end and a slightly negative charge at the other end.

    6. a. very soluble
       d. slightly soluble

    8. b. This is an ionic compound because it has a high melting point, is very soluble, is conductive in solution, and has a hard crystal formation.

       e. This is a molecular compound because it has a low boiling point and is not very soluble in water.

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Lesson 4

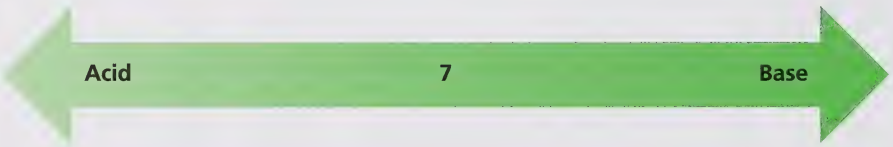
Acids and Bases

When was the last time you put vinegar on your French fries or washed your hands with soap and water? Did you know that you were using **acids** and **bases**? An acid is a compound that dissolves in water to form a solution with a **pH** lower than 7. A base is a compound that dissolves in water to form a solution with a pH greater than 7.

acid: a compound that, when dissolved in water, forms a solution with a pH less than 7

base: a compound that, when dissolved in water, forms a solution with a pH greater than 7

pH: the measure of how acidic or basic a substance is



Acids and bases are present in the foods you eat, the products you use, and the fluids in your body systems. Your stomach, for example, contains strong hydrochloric acid; other organs in your body produce bases to neutralize stomach acid at certain points in the digestive process.

Turn to page 62 of the textbook and read "Acids and Bases in Your Body."

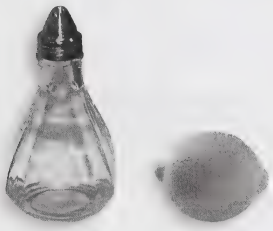
1. Describe two things hydrochloric acid does in your body system.

2. What does sodium hydrogencarbonate do in your body system?

3. What do buffers do?

Check your answers with those on page 100.

You are already familiar with some of the properties of many acids and bases through your own personal contact with things you eat or use in your home. For instance, you probably know that lemons and vinegar taste sour, whereas if you ever had soap in your mouth it tastes bitter. Lemons and vinegar are acids (tastes sour), and soap is a base (tastes bitter). You also know that soap is slippery. This is another property of bases.



Acids

Bases

Turn to page 63 of the textbook and read the information in "Properties of Acids and Bases" and "Indicators."

4. Does vinegar feel slippery?

Never perform a taste or touch test for acids or bases in the laboratory.

5. Describe two tests for acids and bases that are used in the laboratory.

6. What is an indicator?

7. a. What is litmus paper?
   b. What colour does litmus paper turn in an acid and in a base?

8. a. What is a universal indicator?

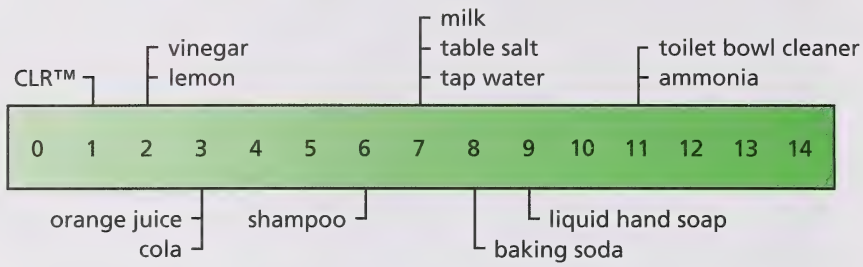
   b. What colour is the universal indicator shown in Figure A2.33 on page 63 of the textbook in acid, base, and neutral solutions?

   c. How does a universal indicator differ from litmus paper?

Check your answers with those on pages 100 and 101.

The pH scale is a measure of the hydrogen ion concentration of a solution. The pH tells you how acidic or basic a substance is. The scale is between 0 and 14 with acidic solutions measuring between 0 and 7 and basic solutions measuring between 7 and 14. A change of a pH of 1 on the scale is a 10× change in acidity of the solution.

pH of Some Common Substances



Turn to page 64 of the textbook and read "The pH Scale." Study Table A2.17.

9. What is the pH of a neutral solution?

10. Draw a line to represent the pH from 0 to 14; then mark the approximate pH of the substances listed on it.

- soap
- blood
- vinegar
- stomach acid
- normal rain water

Check Check your answers with those on page 101.

Acid rain occurs when rain water becomes acidic due to pollutants in the atmosphere. To learn more about acid rain, first read "*info*BIT" on page 64 of the textbook. Then visit the following website:

http://www.scienceman.com/science10

Once there, click on "Unit A: Hot Links" and scroll down to Text Page 64. You will find a number of links to articles and information sites from credible organizations, like Environment Canada and Alberta Environment.

Figure 2.10: Victims of acid rain

Naming Acids

In Lesson 2 of this section you named ionic and molecular compounds using the IUPAC system. The IUPAC system can also be used to name acids. Acids have a hydrogen symbol (H) on the left side of the compound's formula. Organic acids, like CH_3COOH, are the only exception. For these acids, the hydrogen symbol (H) is on the right.

Turn to pages 64 and 65 of the textbook and read "Naming Acids." Examine Table A2.18 closely.

11. What is common in the formulas given in Table A2.18?

12. a. Copy and complete the following table by writing the appropriate IUPAC name and common name for each acid.

Acid	IUPAC Name	Other Name
HCl(aq)		hydrochloric acid
H_2SO_4(aq)	aqueous hydrogen sulfate	
HNO_3(aq)		
H_3PO_4(aq)		

b. How are acids named using the IUPAC system?

13. What are the names of the two organic acids at the bottom of Table A2.18?

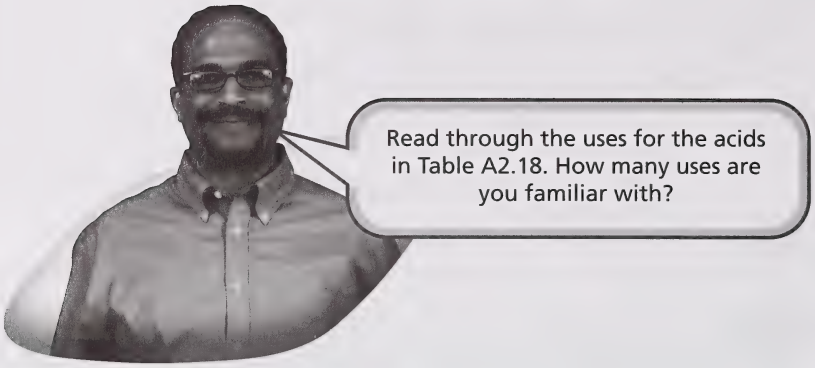
Methanoic acid (formic acid) is part of what gives you the sting when an ant bites you. Methanoic acid is also present in sweat and pine needles.

14. Give the common name for each of the following acids.

a. HCl(aq) b. H_2SO_4(aq) c. CH_3COOH(aq)

15. What are two uses for aqueous hydrogen phosphate, or phosphoric acid?

Check your answers with those on page 101.



Naming Bases

The formulas of some common bases can be recognized by the presence of a hydroxide ion (OH^-) along with a metal ion or ammonium ion.

Turn to page 67 of the textbook and study Table A2.19.

16. What did you notice about the bases that contain a metal ion and a hydroxide ion?

17. Copy and complete the following table by writing the scientific name and the common name for the following bases.

Formula	Scientific Name	Common Name
NaOH(s)		
$Mg(OH)_2(s)$		

18. What are the uses for potassium hydroxide and ammonium hydroxide?

19. You are already familiar with many acids and bases around the house. In a table, list the acids and bases you are familiar with that may be found in a home.

Check your answers with those on pages 101 and 102.

Quicklab

The pH of Common Household Materials

Read the entire activity on page 68 of the textbook.

If you have access to a supervised laboratory, do **Part A**. If you do not have access to a supervised laboratory, do **Part B**.

Part A

Obtain the materials and equipment required to complete this activity. You can test the materials you listed in question 19 if you wish. **Pay special attention to the safety precautions mentioned.**

20. Copy and complete the following table.

Name of Substance	Prediction	pH
lemon juice		
tomato juice		
bleach		
vinegar		
soap		
window cleaner		
drain cleaner		
milk		
baking soda		

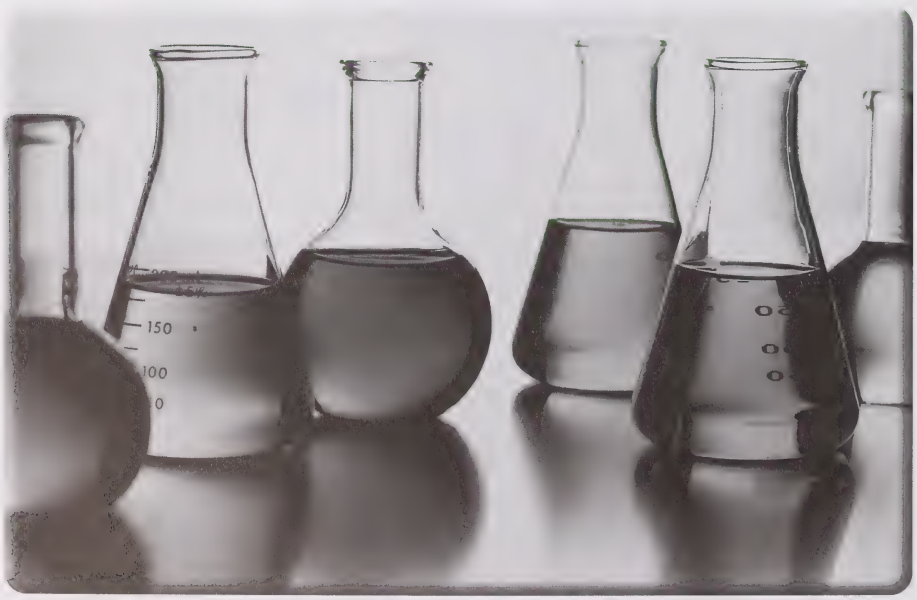
21. Answer questions 1, 2, and 3 of "Questions."

Check your answers with those on pages 102 and 103.

Part B

Use the following table to answer question 21 in Part A.

Name of Substance	Prediction	pH
lemon juice	acid	3
tomato juice	acid	4
bleach	base	11
vinegar	acid	2
soap	base	9
window cleaner	base	10
drain cleaner	base	14
milk	acid	6.5
baking soda	base	8



To see some uses of acids and bases around the home, visit the following website:

www.scienceman.com/science10

Once there, click on "Unit A: Hot Links" and scroll down to Text Page 68. You will find a list of informative Internet links.

Neutralization

What do you think happens when you mix an acid and a base? Turn to page 68 of the textbook and read "Neutralization."

22. What happens to the properties of an acid and a base when the two are mixed together?

23. What are the products produced when an acid and a base are mixed?

Check Check your answers with those on page 103.

Looking Back

You have now completed the concepts for this lesson. You identified some of the properties of acids and bases. You observed the effect of acids and bases on different indicators and defined pH. You identified some ways of naming acids and bases and recognizing their formulas. You compared the pH of some acids and bases in the home and defined neutralization in terms of the properties of acids and bases.

24. Answer questions 1, 3.a., 3.c., 3.d., 4.a., 4.c., 4.h., 6, and 7 of "Check and Reflect" on page 69 of the textbook.

Check Check your answers with those on page 103.

Go to ... Go to page 7 of Assignment Booklet 1B and answer questions 28 to 33.

Glossary

acid: a compound that, when dissolved in water, forms a solution with a pH less than 7

base: a compound that, when dissolved in water, forms a solution with a pH greater than 7

buffer: a substance that keeps the pH of a solution nearly constant even when small amounts of acid or base are added

indicator: a substance that shows whether a solution is acidic or basic

e.g., litmus and red cabbage juice

neutral: having a pH of 7

neutralization: a process in which an acid and a base are mixed to form a solution that loses the acidic and basic properties

pH: the measure of how acidic or basic a substance is

universal indicator: a mixture of several indicators that changes colour in a solution as the acidity of the solution changes

Suggested Answers

1. Hydrochloric acid in your body system dissolves food and activates an enzyme called pepsin.

2. Sodium hydrogencarbonate neutralizes stomach acid, thus deactivating pepsin.

3. Buffers keep the pH of a solution nearly constant despite the addition of a small amount of acid or base to the solution.

4. No, vinegar does not feel slippery.

5. Reaction with metals and pH are two tests used for acids and bases in the laboratory. Acids react with most metals, whereas bases do not. The pH of acids is less than 7, whereas the pH of bases is greater than 7.

6. An indicator is a chemical that will change colour depending on the pH of a solution.

7. a. Litmus paper is a strip of absorbent paper with a chemical, called litmus, dried on it. Litmus paper is used to determine whether a solution is an acid or a base.

   b. Red litmus paper remains red in an acid and turns blue in a base. Blue litmus paper turns red in an acid and remains blue in a base.

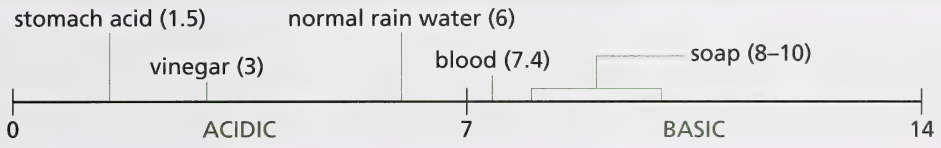
8. a. A universal indicator is a mixture of several indicators that changes colour as the pH of the solution changes.

   b. The universal indicator shown in Figure A2.33 on page 63 of the textbook is pink in an acid and violet in a base. It is yellow in a neutral solution.

   c. Litmus paper only tells you whether a substance is an acid or a base. A universal indicator gives an estimate of the pH of the solution.

9. The pH of a neutral solution is 7.

10.



11. A hydrogen symbol (H) and the subscript (aq) is common to the formulas in Table A2.18.

12. a.

Acid	IUPAC Name	Other Name
HCl(aq)	aqueous hydrogen chloride	hydrochloric acid
H_2SO_4(aq)	aqueous hydrogen sulfate	sulfuric acid
HNO_3(aq)	aqueous hydrogen nitrate	nitric acid
H_3PO_4(aq)	aqueous hydrogen phosphate	phosphoric acid

   b. In the IUPAC system, acids are named as aqueous substances with the name of the compound.

13. Methanoic acid, HCOOH(aq), and ethanoic acid, CH_3COOH(aq), are given at the bottom of the table. You will learn more about naming organic acids if you take a chemistry course in your future years.

14. a. muratic acid b. battery acid c. acetic acid

15. Aqueous hydrogen phosphate, or phosphoric acid, is used to provide the sour taste in soft drinks and in fertilizer manufacturing.

16. Bases that contain a metal ion and a hydroxide ion are all solids.

17.

Formula	Scientific Name	Common Name
NaOH(s)	sodium hydroxide	caustic soda
$Mg(OH)_2(s)$	magnesium hydroxide	milk of magnesia

18. Potassium hydroxide, KOH(s), is used in tanning leather. Ammonium hydroxide, $NH_4OH(aq)$, is used in window cleaner.

19.

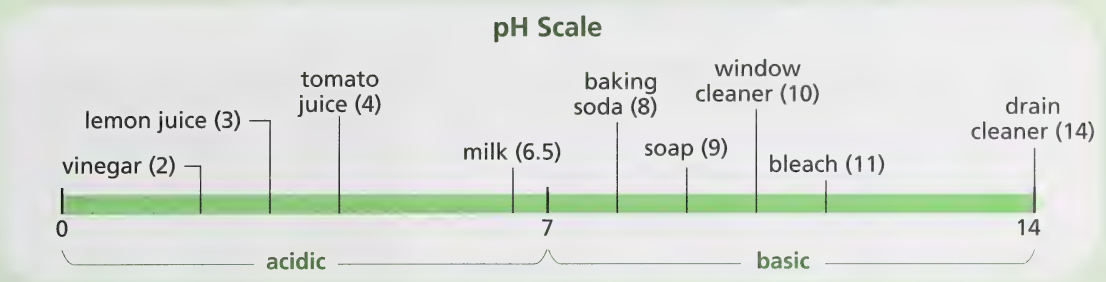
Acids	Bases
• lemon juice • fruit juices, pop • vinegar • some toilet bowl cleaners • products that remove lime (e.g., CLR and Limeaway)	• soap • window cleaner • oven cleaner • drain cleaner • household bleach

20.

Name of Substance	Prediction	pH
lemon juice	acid	3
tomato juice	acid	4
bleach	base	11
vinegar	acid	2
soap	base	9
window cleaner	base	10
drain cleaner	base	14
milk	acid	6.5
baking soda	base	8

21. **Textbook questions 1, 2, and 3 of "Questions," p. 68**

1. to 3. Your completed line, although vertical, should be similar to the following.



22. When an acid and a base are mixed together, the acidic and basic properties disappear.

23. Water and a salt are produced through a chemical process called neutralization.

24. **Textbook questions 1, 3.a., 3.c., 3.d., 4.a., 4.c., 4.h., 6, and 7 of "Check and Reflect," p. 69**

1. a. basic b. acidic c. acidic d. neutral e. acidic
   f. neutral g. acidic h. basic i. neutral j. basic

3. a. base c. neither d. acid

4. a. $HNO_3(aq)$
   c. $CH_3COOH(aq)$
   h. aqueous hydrogen bromide or hydrobromic acid

6. A universal indicator will show what the approximate pH is by changing colour at various pH values. Litmus paper will only turn red or blue to indicate whether the solution is acidic or basic respectively.

7. a. The pH of the solution will increase as more NaOH solution is added.

   b. The conductivity of the solution does not change much as both HCl(aq) and NaOH(aq) are highly conductive as is the NaCl product that is produced.

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Lesson 5

Our Chemical Society

Examine the rooms in your home, and note the items that contain chemicals. Did you notice that pretty much everything contains chemicals? Chemicals are everywhere. Oxygen, water, sand . . . they are all chemicals. These chemicals have been around for billions of years and are part of the natural environment.

Modern society makes thousands of products using chemicals from the natural environment. Many of these products cause harmful effects to the environment, especially when they are not disposed of properly. Products like these include paints, plastics, fertilizers, pesticides, old computer parts, and so on.

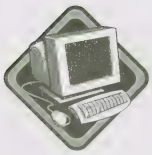
Issues relating to the safe use and disposal of toxic or potentially hazardous compounds are often in the news. Guidelines for the safe use and disposal of the thousands of chemicals used today help protect the environment and life within it.

Turn to page 70 of the textbook and read the introductory paragraphs of "Our Chemical Society." Then read "Issues Related to Chemicals" on pages 70 and 71.

1. Are plastic products toxic?

2. What are the hazards of dumping plastics in landfills?

3. What hazard do mercury batteries present to the environment and life within the environment?



Note: You can research the hazards of mercury in the environment on the Internet. Here is one site you can visit:

http://www.ec.gc.ca/MERCURY/EN/index.cfm

4. Why were chlorofluorocarbons (CFCs) first used in cooling systems?

5. What damage do CFCs cause to the environment?

6. a. What changes have been made over the years to reduce the damage of CFCs to the environment?

   b. What are the results of the changes in the use of CFCs?

Check Check your answers with those on page 108.

For more information on CFCs, visit

http://www.cmdl.noaa.gov/noah/publictn/elkins/cfcs.html

The use of chemical substances for personal use is also an issue as some of these can be toxic or have harmful effects with prolonged use. Two of the main chemical substances that cause harmful effects with prolonged use are alcohol and nicotine. Dozens of other chemicals in tobacco also cause harmful effects with prolonged use.

For more the information on the harmful effects of the prolonged use of alcohol and nicotine, read "Health Concerns" on pages 71 to 73 of the textbook.

7. When does ethanol become toxic to the human body?

8. a. How can the excessive use of alcohol affect the body?

   b. Describe two types of dependence that can occur with excessive use of alcohol.

   c. Describe three ways in which alcohol is a significant killer.

9. a. Name three types of tobacco products that contain nicotine.
   b. How is the use of tobacco products similar to the misuse of alcohol?
   c. Name four body areas where tobacco use causes cancer.
   d. Name five other ailments caused by cigarette smoking.

10. Who are passive smokers and what risks do they have?

11. What steps have been taken in some cities and towns to help passive smokers?

Check Check your answers with those on pages 108 and 109.

Many hazardous chemicals are used in the production of thousands of products that people take for granted. The handling, transportation, and storage of these hazardous chemicals is strictly regulated by government. However, spills and leakages do occur.

Read the information from the middle of page 73 to the middle of page 74.

12. Name a dangerous chemical used in industry and transported in large quantities.

Check Check your answer with the one on page 109.

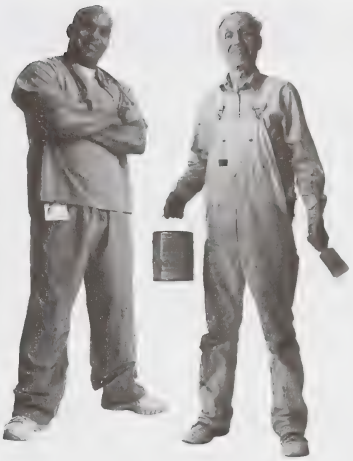
For more information on the hazards of benzene, visit the following website:

http://www.scienceman.com/science10

Once there, click on "Unit A: Hot Links" and scroll down to Text Page 73. You will find links to credible websites that provide ample information.

WWW

What do you think of when you are asked about careers in chemistry? Do you think of a person working in a laboratory or a chemical plant? These are definitely careers in chemistry. However, many other jobs involve the use of chemicals. They include auto mechanics, dentists, hairstylists, mechanics, nurses, and painters.



Read the information in "Chemistry-Related Careers" and "Working with Chemistry" on pages 74 and 75 of the textbook.

13. Research the job of a food technologist. Use your local library or the following website:

    http://jobguide.thegoodguides.com.au/text/jobdetails.cfm?jobid=79

    a. What does a food technologist do?
    b. List four tasks of a food technologist.
    c. Why are food technologists required to wear protective clothing?
    d. Name three personal requirements for the job of food technologist.

Check Check your answers with those on pages 109 and 110.

For more information on a variety of chemistry careers, visit

http://www.chemistrycoach.com/chemistryprofession.htm

Looking Back

You have now completed the concepts for this lesson. You identified some issues related to chemicals in today's society and described some of the health concerns that have surfaced from the use of products containing harmful chemicals. You also identified some issues in the transportation of hazardous chemicals and briefly analyzed some careers connected with chemistry.

14. Answer questions 2, 6, and 8 of "Check and Reflect" on page 75 of the textbook.

Check Check your answers with those on page 110.

Go to ... Go to page 8 of Assignment Booklet 1B and answer questions 34 to 36.

Glossary

addiction: a physical dependence on a substance

nicotine: a highly addictive substance present in tobacco

regulated substance: a substance that must be transported and used according to strict guidelines

Suggested Answers

1. Plastics are not toxic in the final form; but many plastics are produced from toxic compounds.

2. Plastics dumped in landfills take up a lot of space and do not readily break down.

3. Mercury and mercury compounds are extremely toxic. Although the mercury contained in a battery appears safe, it will eventually leak out over the years, especially if the battery is thrown in a landfill or anywhere else in the environment for that matter.

4. Chlorofluorocarbons (CFCs) were first used in cooling systems because they are non-toxic and non-flammable.

5. CFCs contain chlorine, which acts as a catalyst in the upper atmosphere and causes the destruction of Earth's ozone layer.

6. a. CFCs have been replaced with other products and are no longer manufactured in Canada or used in cooling systems. Also, older appliances must have the CFCs in them removed and recycled or destroyed.

   b. Changes in the use of CFCs have led to a reduction of CFC emissions to one-seventh the 1990 value in the 10-year period (from 1990 to 2000). However, since the effects are long lasting, the damage to the ozone layer will not be eliminated for 50 to 60 years.

7. Ethanol becomes toxic to the human body when used in large amounts. Alcohol poisoning can result with extreme alcohol consumption.

8. a. Ways in which excessive alcohol can affect the body are as follows:

   - slurred speech, blurred vision, reduced inhibitions
   - destruction of brain cells
   - increased heart rate and blood pressure
   - chronic high blood pressure
   - loss of body heat due to increased blood flow to skin
   - irritation of linings of stomach and intestines, leading to vomiting
   - the hardening of tissues and destruction of liver cells, leading to cirrhosis of the liver
   - increased urine production

b. Two types of dependence that can occur with excessive alcohol use is physical dependence (alcohol addiction) and psychological dependence. In physical dependence, the body becomes used to the drug and requires it to function. In psychological dependence, the use of the drug is associated with moods or feelings.

c. Ways in which alcohol is a significant killer are as follows:

- Alcohol misuse can lead to high-risk behaviour (e.g., drinking and driving).
- Binge drinking can lead to alcohol poisoning—alcohol levels become so high that it affects the ability of the brain to control a person's breathing and heartbeat.
- A person may also suffocate on vomit if they lose consciousness.
- Mixing alcohol and other drugs, including prescription or over-the-counter drugs, can cause death even at relatively low levels of blood alcohol.

9. a. Cigarettes, cigars, and chewing tobacco contain nicotine.

b. Like the misuse of alcohol, the use of tobacco products causes physical and psychological dependence.

c. Tobacco use causes cancer in the mouth, throat, lungs, and stomach.

d. Five other conditions caused by cigarette smoking are

- nicotine addiction
- cataracts
- wrinkles caused by decreased blood flow to skin
- respiratory diseases (e.g., emphysema)
- circulatory diseases (e.g., high blood pressure and heart disease)

10. Passive smokers are those who are exposed to smoke from other people's tobacco products. Passive smokers are at increased risk of lung cancer; itchy, watery eyes; and throat irritation.

11. Many cities and towns are passing by-laws that prohibit smoking in public places, such as workplaces, restaurants, and bars.

12. Benzene is a hazardous chemical used in industry and transported in large quantities.

13. a. Food technologists develop new food products and improve existing food products. They also set standards for producing, packaging, and marketing food.

b. Food technologists perform a number of tasks, including

- maintaining hygienic conditions during food processing and packaging
- testing raw ingredients and processed food for nutritional value and quality
- developing new products and techniques to produce products
- supervising cleaning and maintenance of equipment used in food processing
- supervising the transportation of foods and inspecting for spoilage
- developing quality-control procedures for product development
- comparing products of various brands
- analyzing results of experimental data
- supervising technicians in the food industry

c. Food technologists are required to wear protective clothing so they don't invalidate tests by contaminating the products they are working with. The protective clothing also protects workers from taking contamination from the job site to their homes.

d. Personal requirements include

- self-initiative
- enjoying scientific activity and working with a high degree of accuracy
- the ability to work as part of a team
- very observant
- physically fit

14. **Textbook questions 2, 6, and 8 of "Check and Reflect," p. 75**

2. Alcohol is considered to be a drug because it affects on the body.

6. Although a product may not be toxic, the chemicals that it is made from may be toxic. Examples include acetate, aspirin, medicines that require benzene in their production, and various plastics (which require acetone during production).

8. Cosmetics formulators need to know the properties of the materials they use and their effects on the different parts of the body. The formulators must be able to perform quality control on the products they produce.

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Conclusion

In this section you reviewed the periodic table and the elements that make up the periodic table. You described the structures of the atoms that make up the different elements and developed an understanding of how atoms combine chemically by gaining, losing, or sharing electrons. You then became familiar with naming substances and categorizing them based on the properties they exhibit. Lastly, you considered how toxic and hazardous chemicals affect the environment and society.

Each person has a responsibility to know the hazards of the chemicals they use and how to use and dispose of these hazardous products safely. You identified a number of occupations that work with and use hazardous chemicals. Many workers—painters, artists, and hair stylists, etc.—use chemicals that may present hazards to health if not properly used.

Chemical Change

Have you seen a rusting vehicle or a rusted metal can? Rust is the result of a chemical reaction between the iron metal, water, and oxygen. A chemical change occurs as atoms combine with other atoms to form a new material called rust.

Chemical reactions occur naturally in various biochemical processes, such as during photosynthesis and cellular respiration. In industry, chemical reactions are used in a variety of applications to produce new materials or store chemical energy.

In this section you will identify a number of chemical reactions and use the law of conservation of mass to help you write balanced chemical equations. You will describe five common types of chemical reactions and write balanced chemical equations for each. You will then define the concept of the mole as a quantity used by chemists to measure the amounts of chemicals in a reaction.

Turn to page 78 of the textbook and read the introduction to Unit A 3.0. Pay particular attention to the key concepts and the learning outcomes listed. They will give you a brief overview of the concepts you will cover in this section.

Lesson 1

Important Examples of Chemical Change

chemical change: the change that occurs when substances react in a chemical reaction to create different substances

When was the last time you watched an awesome display of fireworks? Fireworks is one example of **chemical change** occurring. Other examples of chemical change include dough rising, the changing taste of food cooking on a barbecue, the combustion of fuel in a motor vehicle, a glowing glow stick, and the changing temperature of a hot or cold pack.

Turn to page 79 of the textbook and read the introductory paragraphs of "Important Examples of Chemical Change."

1. Answer questions 1 to 4 of "Minds On . . . Energy Flow Through Systems."

2. Which chemical reaction would be faster—the one that releases the energy of batteries to power a digital camera or the one that recharges the batteries after they have run down?

Check your answers with those on page 120.

In Section 2 you studied evidence of chemical reactions. Evidence of chemical reactions include the formation of gases and precipitates. Familiar examples of chemical reactions that result in the formation of gases are breads and cakes rising and the inflation of an automobile air bag. Some reactions form a solid that hardens over a short period of time. Epoxy glue is one example.

Read "Reactions That Form Gases" and "Reactions That Form Solids" on pages 79 and 80 of the textbook.

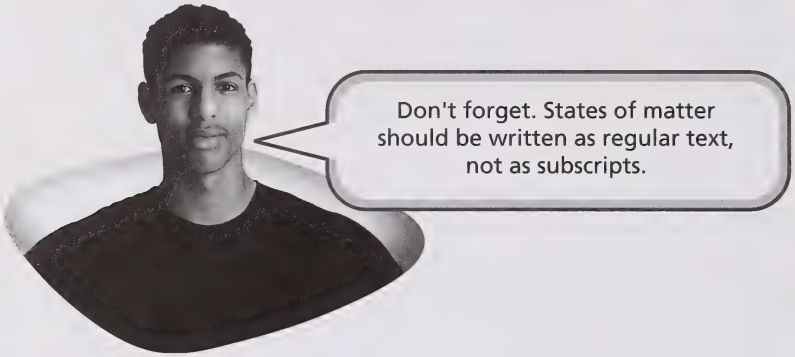
3. How does an automobile airbag inflate?

4. Explain the process of how epoxy glue works.

Check your answers with those on page 120.

Since chemical reactions that start as liquids may result in gases or solids, an important feature of writing chemical equations is to indicate the state of each substance in the chemical reaction. The state of each substance in the chemical reaction is taken at room temperature (25°C) unless otherwise stated.

Read "Showing States in Chemical Formulas" and "*info*BIT" on page 80 of the textbook. You will be provided with guidelines of the state of most elements and many compounds. It will be beneficial for you to remember these guidelines.



5. Write the symbol used for each of the following states.

   a. solid b. liquid c. gas d. aqueous

6. What does *aqueous* mean?

7. Determine the state of each of the following elements and compounds at room temperature. Explain how you decided on the state of each.

   a. Mg b. F_2 c. Br_2 d. K_2SO_4

8. How can the state of some molecular compounds—such as natural gas, gasoline, and beeswax—be used to give a rule for predicting the state of other unfamiliar molecular compounds?

Check your answers with those on page 121.

Exothermic and Endothermic Reactions

In a hot pack and some cold packs, you have to squeeze the package to break a container inside so chemicals can mix and react. If you have used a hot pack in your boots or gloves on a cold, winter day, you know that your feet or hands feel warm. If you have used a cold pack on an injury to help reduce the swelling, you know it feels cold.

exothermic reaction: a reaction that releases energy

endothermic reaction: a reaction that absorbs energy

At the beginning of this lesson, it was stated that energy flow is a part of any chemical reaction. Energy can flow out of a chemical reaction or into a chemical reaction. If energy flows out of a chemical reaction, it is an **exothermic reaction**. If energy flows into a chemical reaction, it is an **endothermic reaction**.

For more information on exothermic and endothermic reactions, turn to page 81 of the textbook and read "Energy Changes."

9. a. Describe what is happening inside a hot pack when you use it.
   b. What type of reaction is occurring?

10. a. Describe what is happening inside a cold pack when you use it.
    b. What type of reaction is occurring?

11. Describe another important exothermic chemical reaction.

12. In a chemical equation, how do you indicate that energy is released in a chemical reaction?

13. If a lead-acid battery can be recharged, why doesn't it last forever?

Check your answers with those on page 121.

biochemical reaction: a reaction that occurs in living things

Two examples of chemical reactions important to life on Earth are photosynthesis and respiration. These two **biochemical reactions** form the basis of life as almost all food produced on Earth begins with photosynthesis.

Turn to page 82 of the textbook and read "Biochemical Reactions."

14. How are photosynthesis and respiration related?

15. Which biochemical reaction is exothermic and which one is endothermic?

Check your answers with those on page 121.

reactant: a substance that reacts in a chemical reaction

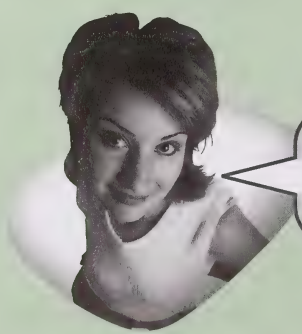
product: a substance that is produced in a chemical reaction

When chemicals react, **reactants** combine and exchange atoms and become different substances, known as **products**. In this exchange of atoms, is energy lost or gained? To find out, you will perform and/or analyze an activity involving chemical reactions.

Inquiry Lab

Mass Change in Chemical Reactions

Read the entire activity on page 83 of the textbook.



If you have access to a supervised laboratory, do **Part A**. If you do not have access to a supervised laboratory, do **Part B**.

Part A

Follow the instructions in Part 1 and Part 2 of the procedure on page 83.

Pay special attention to the safety precautions mentioned. Be sure to wear safety goggles, an apron, and gloves when performing this experiment; and be aware of the WHMIS symbols associated with each chemical.

16. Copy and complete the following table with your observations.

Procedure	Observations	Prediction	Mass Before Reaction	Mass After Reaction
Part 1				
Part 2				

17. Answer the following on page 83 of the textbook.

    a. questions 1, 2, and 3 of "Analyzing and Interpreting"
    b. question 4 of "Forming Conclusions"

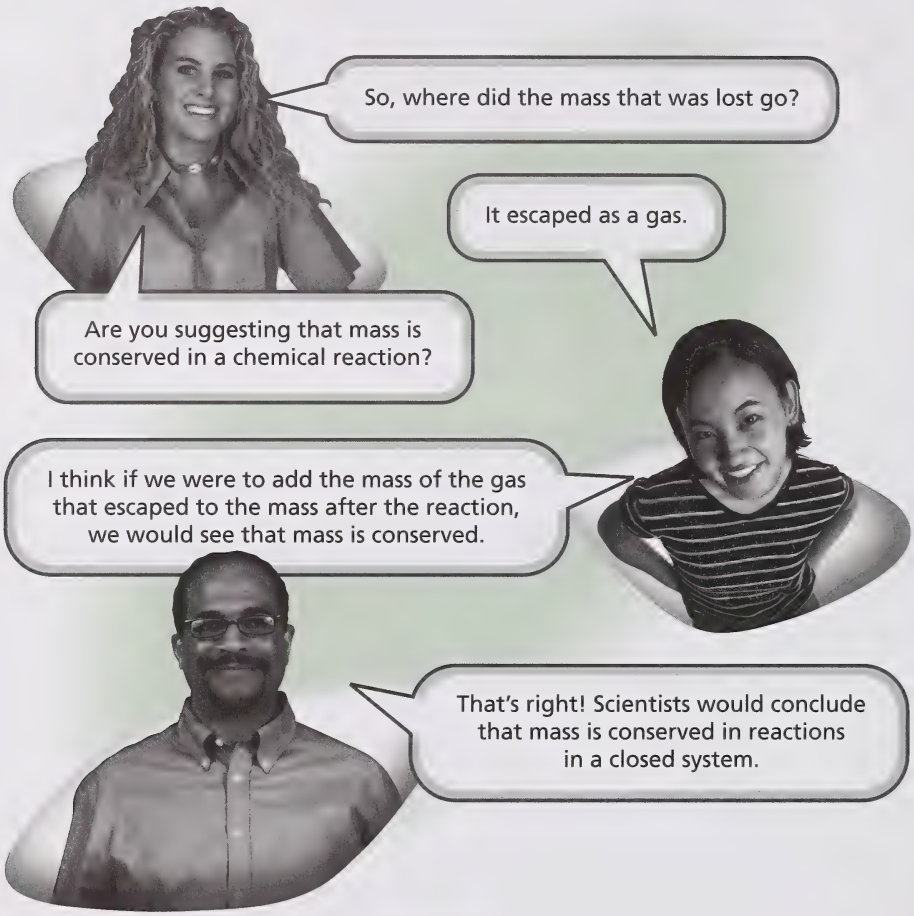
Check your answers with those on page 122.

Part B

Use the following observations to answer question 17 in Part A.

Procedure	Observations	Prediction	Mass Before Reaction	Mass After Reaction
Part 1	white precipitate forms	no change in mass	85.79 g	85.77 g
Part 2	bubbles appear	no change in mass	67.34 g	67.08 g

Insert the Science 10 Multimedia CD into your computer, and view the segment "Chemical Reactions: Generating or Absorbing Heat?" In this segment, two chemical reactions are shown. The initial and final temperatures and masses are recorded for each reaction. Pay particular attention to the change in mass in each reaction. The change in mass in the segment is similar to the change in mass recorded by the student who performed the previous Inquiry Lab.


So, where did the mass that was lost go?
It escaped as a gas.
Are you suggesting that mass is conserved in a chemical reaction?
I think if we were to add the mass of the gas that escaped to the mass after the reaction, we would see that mass is conserved.
That's right! Scientists would conclude that mass is conserved in reactions in a closed system.

18. The activity you just completed suggests that mass is conserved in reactions in a closed system.

    a. What do you think would be the result if you collected the gas that escaped and determined its mass?

    b. Does this suggest that mass is conserved in chemical reactions?

Check your answers with those on page 122.

Turn to pages 84 and 85 of the textbook and read "Conservation of Mass."

19. In the example of burning magnesium in the presence of oxygen, it states that 23.0 g of magnesium metal was recorded before the reaction and 39.0 g of magnesium oxide was recorded after the reaction. Where does the additional 16.0 g come from?

20. What did Antoine Lavoisier propose regarding the mass of reactants before a chemical reaction and the mass of products after the chemical reaction?

21. State the law of conservation of mass in terms of atoms before and after a chemical reaction.

Check Check your answers with those on page 122.

Looking Back

You have just completed the concepts for this lesson. You described a number of examples of chemical change and identified chemical reactions that are significant to society. You also described evidence of chemical change and defined exothermic and endothermic reactions.

At the top of page 84 of the textbook is a summary of the characteristics of chemical reactions. Read this summary to confirm your knowledge of chemical reactions.

22. Answer questions 2, 3, 7, 8, and 11 of "Check and Reflect" on page 85 of the textbook.

Check Check your answers with those on page 123.

Go to ... Go to pages 1 and 2 of Assignment Booklet 1C and answer questions 1 to 4.

Glossary

biochemical reaction: a chemical reaction that occurs in living things

catalyst: a chemical that speeds up a reaction

chemical change: the change that occurs when substances react in a chemical reaction to create different substances

combustion: a chemical reaction that occurs when oxygen reacts rapidly with a substance to form a new substance and release energy

endothermic reaction: a reaction that absorbs energy

exothermic reaction: a reaction that releases energy

product: a substance that is produced by a chemical reaction

reactant: a substance that reacts in a chemical reaction

Suggested Answers

1. **Textbook questions 1 to 4 of "Minds On . . . Energy Flow Through Systems," p. 79**

    1. Fireworks (a) and the glow stick (d) involve a release of energy. Fireworks release heat, light, and sound energy. The glow stick releases light energy.

    2. The boiling water (b) and the baking (c) involve the absorption of energy. Both absorb heat energy produced by a gas flame or an electric current passing through an element.

    3. Chemical change involves a change to the composition of the substance. Therefore, chemical changes occur in (a), (c), and (d).

    4. Physical change involves a change of state (e.g., liquid to gas). A physical change occurs in (b).

2. The reaction that recharges your batteries would be faster. Recharging is done in as little as 30 min. The battery in the camera releases its energy in small amounts over a longer period of time. **Note:** These reactions are controlled to provide the best application for the intended use.

3. An automobile airbag inflates when sodium azide, $NaN_3(s)$, is heated and ignited by an electrical current due to a front-end collision. The chemical reaction results in the rapid release of nitrogen gas that is produced as the $NaN_3(s)$ decomposes.

4. Epoxy glue uses two chemicals that, when mixed together, results in a mixture that begins to harden within five minutes. A molecule of the resin is joined in several places with molecules of hardener, making a web of links that form a large molecule.

5. a. (s) b. (l) c. (g) d. (aq)

6. *Aqueous* means dissolved in water.

7. a. Mg(s)—Magnesium is a metal. All metals, except mercury, are solid at room temperature.

   b. $F_2(g)$—Fluorine is a diatomic element. Most diatomic elements are gases at room temperature.

   c. $Br_2(l)$—Bromine is one of the exceptions to the rule that diatomic elements are gases.

   d. $K_2SO_4(s)$—Potassium sulfate is an ionic compound. All ionic compounds are solid at room temperature.

8. The relative size of the molecules is used to make a general rule for the state of molecular compounds. If the molecule is small, such as natural gas (which has only five atoms), the state of the molecular compound is most likely a gas. If the molecule is moderately large, such as octane (a component of gasoline that contains 20 atoms), the state of the molecular compound is most likely liquid. If the molecule is very large, such as beeswax (which contains 56 atoms), the state of the molecular compound is most likely a solid.

9. a. The chemicals in the hot pack mix and the reaction releases energy in the form of heat.
   b. The reaction is exothermic.

10. a. The chemicals in the cold pack mix and the reaction absorbs energy in the form of heat from its surroundings.

    b. The reaction is endothermic.

11. Another important exothermic reaction is the combustion of a fossil fuel.

12. To indicate that energy is released in a chemical reaction, add "+ energy" on the product side of the chemical equation.

13. A lead-acid battery does not last forever because, eventually, its components, such as the acid and the separator plates, break down.

14. Photosynthesis and respiration are related in that one is the reverse of the other. Photosynthesis uses energy, carbon dioxide, and water to produce glucose and oxygen. Respiration combines glucose and oxygen to produce carbon dioxide and water and release energy.

15. Photosynthesis is endothermic because it uses energy to proceed. Respiration is exothermic because it releases energy.

16. Your completed table should be similar to the following.

Procedure	Observations	Prediction	Mass Before Reaction	Mass After Reaction
Part 1	white precipitate forms	no change in mass	85.79 g	85.77 g
Part 2	bubbles appear	no change in mass	67.34 g	67.08 g

17. a. **Textbook questions 1, 2, and 3 of "Analyzing and Interpreting," p. 83**

    1. In Part 1, a white solid precipitate formed when the liquids were mixed. In Part 2, bubbles appeared when the liquid was mixed with the powder.

    2. In Part 1, mass changed very slightly. The slight change would most likely be due to measurement error. This is the same result as predicted. In Part 2, there was a significant change in mass (a loss). This was not predicted.

    3. Bubbles appeared and then disappeared. This indicated that something was coming out of the solution. Because the flask was open, the gas released was able to escape. The mass lost can be accounted for by the mass of the escaping gas.

    b. **Textbook question 4 of "Forming Conclusions," p. 83**

    4. In a closed system, mass does not change. In an open system, mass may change during a chemical reaction due to escaping gases.

18. a. The mass of the gas would probably be 0.26 g, which is the amount of mass lost in the chemical reaction in Part 2 of the Inquiry Lab.

    b. Yes, this suggests that mass is conserved in chemical reactions.

19. The additional 16.0 g comes from the oxygen consumed from the air during the burning of magnesium metal.

20. Antoine Lavoisier proposed that the mass of the reactants before a chemical reaction is equal to the mass of the products after the chemical reaction.

21. The total number of atoms present before a chemical reaction is equal to the total number of atoms present after the reaction.

22. **Textbook questions 2, 3, 7, 8, and 11 of "Check and Reflect," p. 85**

2. The symbol *(l)* indicates that a substance is a liquid at room temperature. The symbol *(aq)* indicates the substance has been dissolved in water.

3. The combustion of coal and the combustion of hexane are similar in that both reactions produce carbon dioxide and water and release energy. **Note:** The equation of the combustion of coal on page 81 of the textbook shows only carbon and oxygen as reactants and only carbon dioxide as a product. Coal is actually a hydrocarbon (contains carbon and hydrogen). The combustion of hydrocarbons produces carbon dioxide and water.

7. Living systems use enzymes as catalysts to make the reaction proceed at lower temperatures.

8. The chemical reaction would need to occur quickly, be exothermic, and produce large volumes of gases.

11. $\text{mass of hydrogen} = \text{total mass} - (\text{mass of carbon} + \text{mass of oxygen})$
$= 100.0\text{ g} - (40.0\ g + 53.3\text{ g})$
$= 100.0\text{ g} - 93.3\text{ g}$
$= 6.7\text{ g}$

The mass of hydrogen in the sample is 6.7 g.

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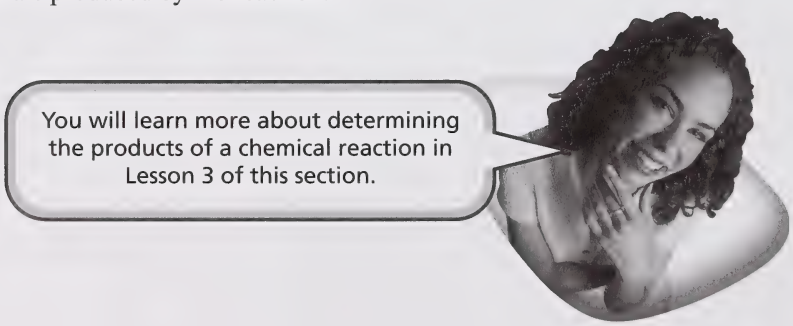
Lesson 2

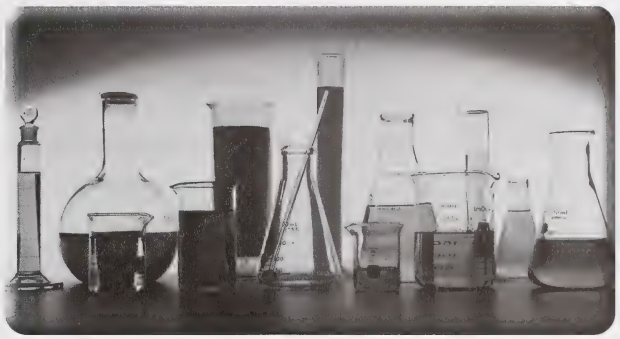
Writing Chemical Equations

Chemical change involves chemical reactions. To record what occurs in a chemical reaction, chemists use a balanced chemical equation made up of chemical formulas.

Before you can write a word equation or a chemical equation, you need to know the reactants and products in the reaction. If you are performing an experiment, you probably know which reactants you are using. You can then determine the products by carefully observing the reaction and by knowing how to analyze the materials produced by the reaction.



Turn to page 86 of the textbook and read the introductory paragraphs of "Writing Chemical Equations" and the information in "Symbolizing Chemical Change." Study the reactions in Figures A3.8 and A3.9 closely.

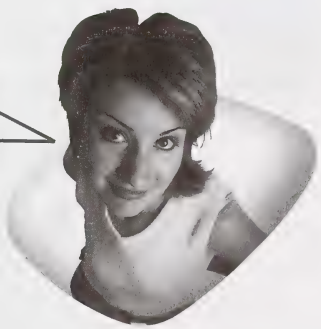


1. How can you recognize that a chemical change has occurred in a chemical reaction?

2. What is the evidence of a chemical change in Figure A3.8?

3. What is the evidence of a chemical change in Figure A3.9?

Check Check your answers with those on page 129.



Read the information in "Writing Word Equations" on pages 86 and 87 of the textbook.

4. Refer to the word equation describing the reaction of magnesium metal with hydrochloric acid on page 86 of the textbook.

   a. How is the word equation for any reaction written?
   b. What is the purpose of the arrow?

5. What happened to the reactants in the reaction of iron and copper(II) chloride?

Check Check your answers with those on page 129.

formula equation: a chemical equation that uses the chemical formulas of the reactants and products to represent a chemical reaction

Writing word equations can become cumbersome for some chemical reactions. So, chemical reactions can be represented by writing a **formula equation**. A formula equation is a chemical equation that uses the chemical formulas of the reactants and products in a chemical reaction. Because mass is conserved in a chemical reaction, you need to write a **balanced formula equation**. A balanced formula equation is a formula equation that has the same number of atoms of each element in both the reactants and products.

Read "Writing Balanced Formula Equations" on pages 87 and 88 of the textbook.

balanced formula equation: a formula equation that has the same number of atoms of each element on both sides of the equation

skeleton equation: a formula equation that only shows the identities of the substances involved in a chemical reaction

Example 3.1

Balance the following **skeleton equation** by adding the proper coefficients in front of each reactant and product so that the total number of atoms of each substance is the same on each side of the equation. **Note:** You cannot change the subscripts.

$$N_2(g) + H_2(g) \rightarrow NH_3(g)$$

Solution

step 1: Count the number of atoms of nitrogen on each side.

There are two atoms of nitrogen on the left side of the equation and only one atom on the right side.

step 2: Balance the nitrogen atoms.

Write the coefficient 2 in front of $NH_3(g)$. This gives two atoms of nitrogen on the left and two atoms of nitrogen on the right.

$$N_2(g) + H_2(g) \rightarrow 2\ NH_3(g)$$

step 3: Count the number of atoms of hydrogen on each side.

There are two atoms of hydrogen on the left and six atoms of hydrogen ($2\ H_3 = 2 \times 3$ or 6) on the right side.

step 4: Balance the hydrogen atoms.

Write 3 in front of $H_2(g)$. This gives six atoms of hydrogen ($3\ H_2 = 3 \times 2$ or 6) on the left side of the equation.

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

The equation is now balanced.

6. How is the formula equation for a reaction related to the word equation?

7. What does a skeleton equation show?

8. Refer to Figure A3.12 on page 87 of the textbook.

   a. How many atoms of hydrogen are on each side of the equation?
   b. How many atoms of oxygen are on each side of the equation?
   c. Is the equation balanced? Explain.

9. When equations you want to balance get more complicated, it is often helpful to use a table to check the number of atoms of each substance on each side. How does Table A3.1 at the bottom of page 88 of the textbook help when balancing the equation for the combustion of methane?

Check your answers with those on page 129.

Balancing chemical equations is used to show the relative proportion of each substance in a chemical equation. It also serves as a reminder that the number of atoms of each kind does not change during a chemical reaction.

Turn to page 89 of the textbook and work through Example Problem A3.1. It provides a step-by-step illustration of balancing a more complicated chemical equation. This method involves using a table to record the number of atoms of each type on both sides of the equation.

10. Refer to Example Problem A3.1.

    a. Why is a 3 written in front of the $Fe(NO_3)_2(aq)$?
    b. Why is a 6 written in front of $NaNO_3(aq)$?
    c. Why is a 2 written in front of $Na_3PO_4(aq)$?

11. Answer questions 1.a., 1.b, and 1.c. of "Practice Problem" on page 89 of the textbook. Use a table like the following to answer and check your answers.

Atom or Polyatomic Ion	Number of Reactant Atoms Before Balancing	Number of Product Atoms Before Balancing	Number of Reactant Atoms After Balancing	Number of Product Atoms After Balancing

Check your answers with those on pages 129 to 132.

Insert the Science 10 Multimedia CD into your computer, and view the flash segment "A Chemical Reaction." Observe what happens when two hydrogen molecules combine with an oxygen molecule.

12. a. How many atoms of hydrogen are on each side of the equation?

    b. How many atoms of oxygen are on each side of the equation?

    c. Is the equation balanced? Why?

Check Check your answers with those on page 133.

Looking Back

You have just covered all the concepts for this lesson. You wrote word equations and skeleton equations given the reactants and products for a chemical reaction. You then balanced these equations to show the correct proportions of each substance in the chemical reaction.

13. Answer questions 6, 7.b., 7.c., 7.d., 8.c., 8.e., 8.g., 9.a., and 9.d. of "Check and Reflect" on page 90 of the textbook.

Check Check your answers with those on pages 133 to 135.

Go to ... Go to pages 2 and 3 of Assignment Booklet 1C and answer questions 5 to 10.

Glossary

balanced formula equation: a formula equation that has the same number of atoms of each element on both sides of the equation

formula equation: a chemical equation that uses the chemical formulas of the reactants and products to represent a chemical reaction

skeleton equation: a formula equation that only shows the identities of the substances involved in a chemical reaction

Suggested Answers

1. You can recognize that a chemical change has occurred in a chemical reaction by carefully observing the changes that occur as the chemical reaction proceeds. Chemical changes may include colour change, the formation of bubbles, the formation of a coloured gas, the formation of a precipitate, and the release or absorption of energy.

2. The visible evidence of a chemical change in Figure A3.8 is the formation of bubbles, the formation of a yellow-brown gas, and a change in colour of the solution to green or blue. Additional evidence includes an increase in temperature and the corrosion of the copper.

3. The visible evidence of chemical change in Figure A3.9 is the formation of bubbles and the slight corrosion (blackening) of the magnesium metal. Additional evidence includes an increase in temperature.

4. a. The word equation for any reaction is written as follows:

   step 1: Write the names of the reactants using a plus sign (+) between each of the reactants.
   step 2: Use an arrow to separate the reactants from the products.
   step 3: Write the names of the products with a plus sign (+) between each of the products.

   b. The purpose of the arrow is to separate the reactants from the products. **Note:** In future chemistry courses, you will learn that the arrow also shows the direction of the reaction.

5. The copper in copper(II) chloride was replaced by the iron to produce iron(II) chloride. **Note:** You will learn more about types of reactions in Lesson 3 of this section.

6. The formula equation uses the formula of each substance named in the word equation.

7. A skeleton equation only shows the identities of all the substances involved in the reaction.

8. a. There are four atoms of hydrogen on each side of the equation.

   b. There are two atoms of oxygen on each side of the equation.

   c. Yes, the equation is balanced because there are the same number of atoms of each substance on both sides of the equation.

9. The table helps check if a formula equation is balanced by showing the number of reactant atoms and product atoms of each element.

10. a. A 3 is written in front of $Fe(NO_3)_2(aq)$ to balance the number of Fe atoms. This will give three Fe atoms on both sides of the equation.

    b. A 6 is written in front of $NaNO_3(aq)$ to balance the number of NO_3^- ions. This will give six NO_3^- ions on both sides of the equation. **Note:** The coefficient 3 and the subscript 2 means there are six (3×2) NO_3^- ions in 3 $Fe(NO_3)_2(aq)$.

c. A 2 is written in front of $Na_3PO_4(aq)$ to balance the number of Na atoms and PO_4^{3-} ions. This will give six Na atoms and two PO_4^{3-} ions on both sides of the equation.

11. **Textbook questions 1.a., 1.b., and 1.c. of "Practice Problem," p. 89**

1. a. $N_2(g) + H_2(g) \rightarrow NH_3(g)$

step 1: Count the number of atoms of each element in the reactants and in the products. Record the numbers in the table.

Atom or Polyatomic Ion	Number of Reactant Atoms Before Balancing	Number of Product Atoms Before Balancing	Number of Reactant Atoms After Balancing	Number of Product Atoms After Balancing
N	2	1		
H	2	3		

step 2: Balance the number of nitrogen (N) atoms.

Because there are two N atoms on the reactant side and only one N atom on the product side, write a 2 in front of $NH_3(g)$.

$$N_2(g) + H_2(g) \rightarrow 2\ NH_3(g)$$

step 3: Balance the number of hydrogen (H) atoms.

With the coefficient 2 in front of $NH_3(g)$, there are six H atoms (2×3) on the product side of the equation. Because there are only two H atoms on the reactant side of the equation, write a 3 in front of $H_2(g)$.

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

step 4: Complete the table to check if the equation is balanced.

Atom or Polyatomic Ion	Number of Reactant Atoms Before Balancing	Number of Product Atoms Before Balancing	Number of Reactant Atoms After Balancing	Number of Product Atoms After Balancing
N	2	1	2	2
H	2	3	6	6

The equation is balanced.

b. $CaC_2(s) + H_2O(l) \rightarrow Ca(OH)_2(s) + C_2H_2(g)$

Atom or Polyatomic Ion	Number of Reactant Atoms Before Balancing	Number of Product Atoms Before Balancing	Number of Reactant Atoms After Balancing	Number of Product Atoms After Balancing
Ca	1	1		
C	2	2		
H	2	4		
O	1	2		

According to the table, there are an equal number of calcium (Ca) and carbon (C) atoms on both sides of the equation. However, there are twice as many hydrogen (H) and oxygen (O) atoms on the product side than there are on the reactant side of the equation. Therefore, write a 2 in front of $H_2O(l)$ to balance the equation.

$$CaC_2(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(s) + C_2H_2(g)$$

Check to see if the equation is balanced.

Atom or Polyatomic Ion	Number of Reactant Atoms Before Balancing	Number of Product Atoms Before Balancing	Number of Reactant Atoms After Balancing	Number of Product Atoms After Balancing
Ca	1	1	1	1
C	2	2	2	2
H	2	4	4	4
O	1	2	2	2

The equation is balanced.

c. $SiCl_4(s) + H_2O(l) \rightarrow SiO_2(s) + HCl(aq)$

Atom or Polyatomic Ion	Number of Reactant Atoms Before Balancing	Number of Product Atoms Before Balancing	Number of Reactant Atoms After Balancing	Number of Product Atoms After Balancing
Si	1	1		
Cl	4	1		
H	2	1		
O	1	2		

There are an equal number of silicon (Si) atoms on both sides of the equation.

Because there are four chlorine (Cl) atoms on the reactant side and only one on the products side, write a 4 in front of HCl(aq).

$SiCl_4(s) + H_2O(l) \rightarrow SiO_2(s) + 4\ HCl(aq)$

With the coefficient 4 in front of HCl(aq), there are four hydrogen (H) atoms (4×1) on the product side of the equation. Because there are only 2 H atoms on the reactant side of the equation, write a 2 in front of $H_2O(l)$.

$SiCl_4(s) + 2\ H_2O(l) \rightarrow SiO_2(s) + 4\ HCl(aq)$

With the coefficient 2 in front of $H_2O(l)$, there are two oxygen (O) atoms (2×1) on the reactant side. This matches the two O atoms on the product side of the equation.

Check if the equation is now balanced.

Atom or Polyatomic Ion	Number of Reactant Atoms Before Balancing	Number of Product Atoms Before Balancing	Number of Reactant Atoms After Balancing	Number of Product Atoms After Balancing
Si	1	1	1	1
Cl	4	1	4	4
H	2	1	4	4
O	1	2	2	2

The equation is balanced.

12. a. There are 4 atoms of hydrogen on each side of the equation.

    b. There are 2 atoms of oxygen on each side of the equation.

    c. Yes, the equation is balanced. There are the same number of atoms of each element on both sides of the equation.

13. **Textbook questions 6, 7.b., 7.c., 7.d., 8.c., 8.e., 8.g., 9.a., and 9.d. of "Check and Reflect," p. 90**

    6. The reactants are the elements or compounds that appear on the left side of the equation. The products are the elements or compounds that appear on the right side of the equation. The reactants are Zn(s) and HCl(aq), and the products are $ZnCl_2$(aq) and H_2(g).

       The state symbols appear immediately following the chemical formulas. The symbol *(s)* indicates a solid; the symbol *(aq)* indicates an aqueous solution; and the symbol *(g)* indicates a gas.

       The formulas are the combinations of letters and numbers representing the elements or compounds. Zn(s), HCl(aq), $ZnCl_2$(aq), and H_2(g) are all formulas in this chemical equation.

       The coefficients are the numbers in front of the formulas. The coefficient 2 in front of HCl, for example, means there are two molecules of hydrochloric acid on the reactants side.

    7. b. $K(s) + O_2(g) \rightarrow K_2O(s)$

       Starting with the oxygen (O) atoms, there are two O atoms on the reactants side and only one O atom on the products side. Write a 2 in front of K_2O(s).

       $K(s) + O_2(g) \rightarrow 2\ K_2O(s)$

       Now, balance the potassium (K) atoms. With the coefficient 2 in front of K_2O(s), there are four K atoms on the products side. Therefore, write a 4 in front of K(s) on the reactants side to balance the equation.

       $4\ K(s) + O_2(g) \rightarrow 2\ K_2O(s)$

       The equation is balanced.

c. $C_6H_{12}O_6(s) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Begin by balancing the carbon (C) atoms. Write a 6 in front of the $CO_2(g)$ on the products side.

$$C_6H_{12}O_6(s) + O_2(g) \rightarrow 6\ CO_2(g) + H_2O(g)$$

Next, balance the hydrogen (H) atoms. Write a 6 in front of the $H_2O(g)$ on the products side.

$$C_6H_{12}O_6(s) + O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(g)$$

Finally, balance the oxygen (O) atoms. Write a 6 in front of $O_2(g)$ on the reactants side.

$$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(g)$$

The equation is balanced.

d. $H_2SO_4(aq) + NaOH(s) \rightarrow Na_2SO_4(aq) + H_2O(l)$

Begin by balancing the sodium (Na) atoms. Write a 2 in front of NaOH(s) on the reactants side.

$$H_2SO_4(aq) + 2\ NaOH(s) \rightarrow Na_2SO_4(aq) + H_2O(l)$$

Check the number of atoms in the remaining elements. There are four hydrogen atoms on the reactants side and only two hydrogen atoms on the products side. Therefore, write a 2 in front of the $H_2O(l)$ on the products side.

$$H_2SO_4(aq) + 2\ NaOH(s) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$$

Count the number of atoms of each element on the reactants side and on the products side. There are the same number of atoms of each element on both sides of the equation. Therefore, the equation is balanced.

8. c. **Skeleton equation:** $Ca(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow NaNO_3(aq) + CaSO_4(s)$

Balanced equation: $Ca(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow 2\ NaNO_3(aq) + CaSO_4(s)$

e. **Skeleton equation:** $Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$

Balanced equation: $2\ Na(s) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + H_2(g)$

g. **Skeleton equation:** $S_8(s) + O_2(g) \rightarrow SO_2(g)$

Balanced equation: $S_8(s) + 8\ O_2(g) \rightarrow 8\ SO_2(g)$

9. a. **Balanced equation:** $Ca(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$

   d. **Balanced equation:** $2\ NO_2(g) \rightarrow N_2O_4(g)$

   Note: "Reacts with itself" simply means nitrogen dioxide molecules combine together.

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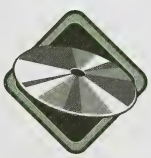
Lesson 3

Five Common Types of Chemical Reactions

When sodium burns in chlorine gas, the compound sodium chloride forms. Two elements combine to form a compound. When an electric current is passed through water, hydrogen gas and oxygen gas form. In this case, a compound is separated into two elements. Other reactions involve an exchange of elements and/or polyatomic ions in two compounds to form two new compounds.

There are millions of chemical compounds and several chemical reactions that can occur for each compound. How do chemists work with all these chemical reactions?

Chemists have found that many chemical reactions have common characteristics. They have been able to group chemical reactions into a number of different types. In this lesson you will examine five common types of chemical reactions.



Insert the Science 10 Multimedia CD into your computer. In the Photo Gallery for Module 1, you can see the results of a couple of chemical reactions.

Turn to page 91 of the textbook and read the introductory paragraphs of "Five Common Types of Chemical Reactions."

1. List the five common types of chemical reactions.

Check your answer with the one on page 147.

Formation Reactions

formation reaction: a chemical reaction in which two elements combine to form a compound

Formation reactions are the simplest type of chemical reaction. The simplest formation reaction occurs when two elements combine to form one compound.

The general form of a formation reaction is

$$A + B \rightarrow AB$$

Note: Chemists use the uppercase letters A, B, C, and D to represent elements in the five common types of reactions.

In this reaction, A and B represent different elements, and AB represents the compound formed.

Turn to pages 91 to 93 of the textbook and read "Formation Reactions." Work through Example Problems A3.2, A3.3, and A3.4 carefully.

2. What is another name for a formation reaction?

3. a. Sulfur reacts with oxygen to produce sulfur dioxide. Write the skeleton equation for this reaction.

   b. Does this equation fit the general pattern for a formation reaction? Explain.

4. Why is the formation of sulfur dioxide of interest to all scientists?

5. a. Name the two elements that combine to form ammonia.
   b. What are two uses of ammonia?

6. What is the purpose of the catalyst used in the synthesis of ammonia? Refer to "*re*SEARCH" at the top of page 89 of the textbook.

7. a. How are the skeleton equations on page 92 of the textbook different from those on page 91?

   b. How do the skeleton equations on pages 91 and 92 fit the form of a formation reaction?

8. Answer questions 2, 3, 4.a., and 4.b. of "Practice Problems" on pages 92 and 93 of the textbook. Refer to Example Problems A3.2, A3.3, and A3.4 if needed.

9. Simple formation reactions involve the combination of elements to form a compound. What do more complex formation reactions involve?

10. Answer questions 1.a., 1.c., 2.b., 2.d., 3.b., 3.c., and 3.f. of "Skill Practice: Formation Reactions" on page 93 of the textbook.

Check your answers with those on pages 147 and 148.

Decomposition Reactions

decomposition reaction: a chemical reaction in which a compound is broken up into its elements

A second common type of chemical reaction is the **decomposition reaction**. A decomposition reaction is the reverse of a formation reaction; however, a decomposition reaction does not proceed by itself. Usually special equipment or energy is required to make a decomposition reaction proceed.

The general form of a decomposition reaction is

$$AB \rightarrow A + B$$

AB represents a compound, and A and B represent different elements. Notice that the general form of a decomposition reaction is the reverse of the general form of a formation reaction.

Turn to page 94 of the textbook and read "Decomposition Reactions." Work through Example Problem A3.5 carefully.

11. How many reactants are there in a decomposition reaction?

12. a. Name the two products in the decomposition of water.

    b. What source of energy is required to make the decomposition of water proceed?

    c. What is the name of the process for the decomposition of water?

13. Answer questions 5.a. and 5.d. of "Practice Problem" on page 94 of the textbook.

Check your answers with those on page 148.

Hydrocarbon Combustion Reactions

hydrocarbon combustion reaction: a chemical reaction in which a hydrocarbon reacts with oxygen to form carbon dioxide and water

hydrocarbon: a compound that contains hydrogen and oxygen

Another common type of chemical reaction is the **hydrocarbon combustion reaction**. **Hydrocarbons** are compounds that contain hydrogen and carbon. A combustion reaction is a reaction with oxygen that is fast and exothermic (releases energy). The general form of a hydrocarbon combustion reaction is

$$C_xH_y + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

Subscripts x and y represent the number of carbon and hydrogen atoms respectively in the hydrocarbon. For example, in the chemical formula for gasoline, C_6H_{14}, x is 6 and y is 14.

Turn to page 95 of the textbook and read "Hydrocarbon Combustion." Work through Example Problem A3.6 closely.

14. Name four examples of hydrocarbons.

15. The chemical formula for methane is CH_4. What is x and y in this formula?

16. a. What are the products of the complete combustion of a hydrocarbon?
    b. What is required for the complete combustion of a hydrocarbon?
    c. Write the skeleton equation for the complete combustion of methane.

17. What does Figure A3.19 show about hydrocarbon combustion reaction involving an oxy-acetylene torch?

18. Answer questions 6.b. and 6.c. of "Practice Problem" on page 95 of the textbook. Refer to Example Problem A3.6 if needed.

Check your answers with those on pages 148 and 149.

Single Replacement Reactions

single replacement reaction: a chemical reaction in which a compound reacts with an element to form a new compound

In many reactions, a metal element reacts with an ionic compound. The metal element replaces the metal ion in the compound and the metal ion frees itself from the compound to end up by itself as a metal element. Because one element is replaced by another, this type of reaction is called a **single replacement reaction**. The general form of a single replacement reaction is

$$A + BC \rightarrow B + AC$$

A is one metal element, B is another metal element, and C is a non-metal element.

Turn to pages 96 and 97 of the textbook and read "Single Replacement Reactions." Work through Example Problems A3.7 and A3.8 closely.

19. a. Which element is replaced when magnesium reacts with silver nitrate?

    b. Why are two silver atoms, Ag(s), replaced by one atom of magnesium, Mg(s)? **Hint:** What is the charge of each ion?

20. In Example Problem A3.7, why are three atoms of copper replaced by two atoms of aluminium?

21. Answer question 8 of "Practice Problems" on page 97 of the textbook.

Check your answers with those on page 149.

Not all single replacement reactions involve the replacement of a metal ion by a metal element. In some single replacement reactions, the non-metal ion is replaced by the non-metal element. (The non-metal ions trade places.) The general form in this case is

$$D + BC \rightarrow C + BD$$

In this formula, D and C are two non-metals and B is a metal.

22. Review Example Problem A3.8 on page 97 of the textbook.

    a. Which two elements exchange in the reaction of liquid bromine and aqueous iron(III) iodide? Are these elements metals or non-metals?

    b. How many atoms of bromine are required to replace the iodine in each formula unit of aqueous iron(III) iodide? How many iodide ions were replaced?

23. Answer question 7 of "Practice Problems" on page 97 of the textbook.

Question 24 will provide you with some additional practice in predicting products and writing equations for decomposition and single replacement reactions.

24. Answer questions 1.a., 1.e., 2.c., 2.f., 3.b., and 3.d. of "Skill Practice: Decomposition and Single Replacement Reactions" on page 97 of the textbook.

Check your answers with those on page 149.

Inquiry Lab

Formation, Decomposition, and Single Replacement Reactions

Read the entire activity on pages 98 and 99 of the textbook. Review the list of materials and equipment and the steps required in each part of the procedure.

If you have access to a supervised laboratory, do **Part A**. If you do not have access to a supervised laboratory, do **Part B.**

Part A

Obtain the required materials and equipment listed on page 98. Complete the four parts as outlined in the procedure on pages 98 and 99. **Pay particular attention to the safety precautions mentioned and the hazards noted for each chemical.**

Follow these additional safety precautions:

- Anhydrous copper(II) sulfate and copper(II) sulfate pentahydrate are both poisonous. Wash your hands thoroughly with soap and water after completing the activity.
- Hydrogen peroxide (3%) is sufficiently diluted, so it will not harm your skin within a few minutes. However, wash immediately if it does contact your skin. Also, use adequate eye protection.
- Use caution when working around Bunsen burners.

25. Answer the following on page 99 of the textbook.

    a. questions 1 to 8 of "Analyzing and Interpreting"
    b. question 9 of "Forming Conclusions"

Check your answers with those on page 150.

Part B

Insert the Science 10 Multimedia CD into your computer, and view the segment "Formation, Decomposition, and Single Replacement Reactions." Use this segment to answer question 25 in Part A. You may need to review parts of the segment more than once.

Double Replacement Reactions

double replacement reaction: a chemical reaction in which two compounds react to form two new compounds.

In a single replacement reaction, a reactive element—metal or non-metal—reacts with an ionic compound. A more complicated replacement reaction occurs when two ionic compounds react to form two new compounds. This is called a **double replacement reaction**. Since ionic compounds are solids at room temperature, these reactions occur when the ionic compounds are in aqueous solutions. This allows the ions to mix. The general form of a double replacement reaction is

$$AB + CD \rightarrow AD + CB$$

In this formula, A and C are positive ions and B and D are negative ions.

Turn to pages 100 and 101 of the textbook and read "Double Replacement Reactions." Work through Example Problem A3.9 carefully.

26. What evidence of a reaction often occurs in a double replacement reaction with ionic compounds?

27. Can positive ions pair up with positive ions and negative ion with negative ions?

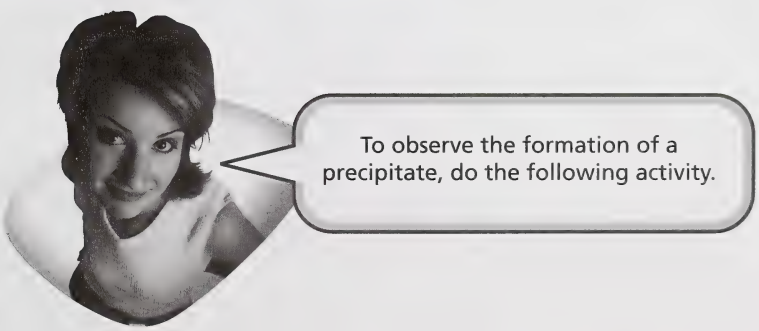
28. Which ions appear first in the chemical formula of the compounds in a double replacement reaction?

29. a. In Example Problem A3.9, which compound formed a precipitate? Why?

    b. Why didn't sodium nitrate, $NaNO_3(aq)$, form a precipitate?

    c. Do these results agree with the information given in Table A2.14 on page 57 of the textbook? Explain.

Check your answers with those on page 151.



Quicklab

Double Replacement Reactions

Read the entire activity on page 101 of the textbook. Review the list of materials and equipment, and pay close attention to the steps outlined in the procedure.

30. What is the purpose of this activity?

Check your answer with the one on page 151.

If you have access to a supervised laboratory, do **Part A**. If you do not have access to a supervised laboratory, do **Part B**.

Part A

Obtain the required materials and equipment listed on page 101 of the textbook, and follow the procedure as outlined. **Pay special attention to the safety precautions mentioned and the hazards noted for each chemical.**

31. Copy and complete the following table. Write *precipitate* and its colour (e.g., *yellow precipitate*) if a precipitate forms when the chemicals are mixed. Write *no precipitate* if a precipitate does not form when the chemicals are mixed.

	Silver Nitrate	Sodium Hydroxide	Calcium Chloride	Sodium Carbonate
Sodium Iodide				
Iron(III) Chloride				
Sodium Carbonate				
Silver Nitrate				

32. Answer question 1 of "Questions."

Check Check your answers with those on pages 151 and 152.

Part B

Insert the Science 10 Multimedia CD into your computer, and view the segment "Double Replacement Reactions." Use the information in this segment to answer questions 31 and 32 in Part A.

You have studied five common types of reactions up to this point in this lesson. Complete question 33 to summarize these five types of reactions.

33. Copy and complete the following table. Insert all five types of reactions and the general statement for each. The formation reaction has been done for you.

Type of Reaction	Reactants	Products
formation (synthesis)	A + B	AB
decomposition		
hydrocarbon combustion		
single replacement		
double replacement		

Check Check your answer with the one on page 152.

Work through a) and b) of Example Problem A3.10 on pages 102 and 103 of the textbook. Use the table you created or Table A3.3 on page 102 of the textbook to help you follow the various types of questions in this example problem.

34. Refer to problem a) of Example Problem A3.10.

    a. Why was the reaction type predicted to be a decomposition reaction?

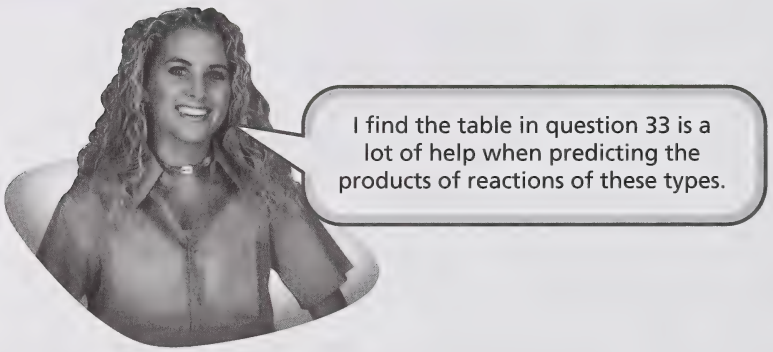
    b. On the products side of the reaction, how can you predict that the lead will be solid and the bromine liquid?

35. Refer to problem b) of Example Problem A3.10.

    a. Why was the reaction type predicted to be a double replacement reaction?

    b. How can you predict that the products will be nickel(III) sulfite and sodium nitrate?

Check Check your answers with those on pages 152 and 153.



Review the remaining parts of Example Problem A3.10 to help you answer the next question. Notice the release of energy shown in Figure A3.22. This release of energy is typical of many formation reactions.

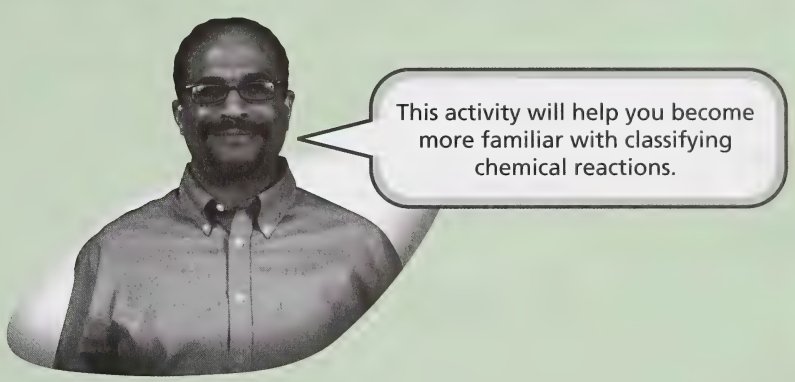
36. Answer questions 10.a., 10.c., and 10.d. of "Practice Problem" on page 103 of the textbook.

Check Check your answers with those on page 153.

Quicklab

Classifying Chemical Reactions

Turn to page 104 of the textbook and read the entire activity. Review the list of materials and equipment, and pay attention to the steps outlined in the procedure.



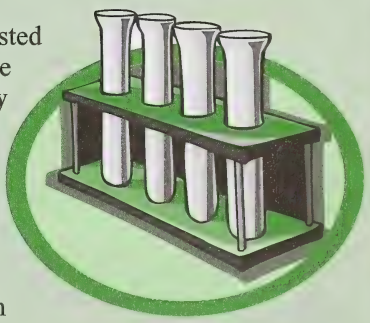
37. What is the purpose of this activity?

Check your answer with the one on page 153.

If you have access to a supervised laboratory, do **Part A**. If you do not have access to a supervised laboratory, do **Part B**.

Part A

Obtain the required materials and equipment listed on page 104, and carefully follow the procedure as outlined. **Pay special attention to the safety precautions mentioned and the hazards noted for each chemical.**



38. Describe what happens in steps 1 to 4 of the procedure.

39. Answer questions 1 to 5 of "Questions" on page 104.

Check your answers with those on pages 153 and 154.

Part B

Insert the Science 10 Multimedia CD into your computer, and view the segment "Classifying Chemical Reactions." Use the information in this segment to answer questions 38 and 39 in Part A.

40. Explain why you should never point the open end of a test tube toward someone while heating the substance in it. (Refer to step 2 of the previous activity.)

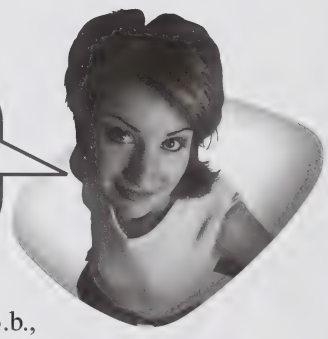
41. Work through Example Problem A3.11 on pages 104 and 105 of the textbook; then answer question 11 of "Practice Problems" on page 105.

Check Check your answers with those on page 154.

Use the Internet to see how ammonia is produced and used for the production of fertilizers. Either enter *fertilizer production* in any of the Internet's search engines or go to the site given in "*re*SEARCH" on page 105 of the textbook.

Looking Back

You have now completed the concepts for this lesson. You identified five types of chemical reactions and learned to write the word equation and balanced formula equation for each type of reaction.



42. Answer questions 1.b., 1.e., 1.g., 1.j., 2.c., 2.g., 3.b., 3.d, 5, and 6 of "Check and Reflect" on page 106 of the textbook.

Check Check your answers with those on pages 154 and 155.

Go to ... Go to pages 4 and 5 of Assignment Booklet 1C and answer questions 11 to 18.

Glossary

decomposition reaction: a chemical reaction in which a compound is broken up into its elements

double replacement reaction: a chemical reaction in which two compounds react to form two new compounds

formation reaction: a chemical reaction in which two elements combine to form a compound

hydrocarbon: a compound that contains hydrogen and carbon

hydrocarbon combustion reaction: a chemical reaction in which a hydrocarbon reacts with oxygen to form carbon dioxide and water

single replacement reaction: a chemical reaction in which a compound reacts with an element to form a new compound

Suggested Answers

1. The five common types of chemical reactions are formation reactions, decomposition reactions, hydrocarbon combustion reactions, single replacement reactions, and double replacement reactions.

2. A formation reaction is also known as a synthesis reaction.

3. a. $S_8(s) + O_2(g) \rightarrow SO_2(g)$
   b. Yes, this equation has two elements that combine to form one compound.

4. The formation of sulfur dioxide is of interest to all scientists because sulfur dioxide can combine with water to form acid rain. Acid rain has serious effects on the environment.

5. a. Nitrogen and hydrogen combine to form ammonia.
   b. Ammonia is used in the production of explosives and fertilizers.

6. The purpose of the catalyst is to speed up the reaction.

7. a. The skeleton equations on page 92 of the textbook all involve the combining of a metal and a non-metal. The skeleton equations on page 91 involve the combination of two non-metals.

   b. The skeleton equations on pages 91 and 92 of the textbook fit the form of a formation reaction because they involve the combination of two elements forming a single compound.

8. **Textbook questions 2, 3, 4.a., and 4.b. of "Practice Problems," pp. 92 and 93**

   2. **Skeleton equation:** $Li(s) + O_2(g) \rightarrow Li_2O(s)$

      Balanced equation: $4\ Li(s) + O_2(g) \rightarrow 2\ Li_2O(s)$

   3. **Skeleton equation:** $Pb(s) + Br_2(g) \rightarrow PbBr_4(s)$

      Balanced equation: $Pb(s) + 2\ Br_2(g) \rightarrow PbBr_4(s)$

   4. a. calcium nitride, $Ca_3N_2(s)$ b. silver oxide, $Ag_2O(s)$

9. More complex formation reactions involve compounds that react to form a single compound. The compounds formed in complex formation reactions are difficult to predict and are left for future chemistry courses.

10. **Textbook questions 1.a., 1.c., 2.b., 2.d., 3.b., 3.c., and 3.f. of "Skill Practice: Formation Reactions," p. 93**

    1. a. potassium iodide c. cesium chloride

    2. b. $Mg(s) + F_2(g) \rightarrow MgF_2(s)$ d. $K(s) + N_2(g) \rightarrow K_3N(s)$

    3. b. $2\ Al(s) + 3\ Br_2(l) \rightarrow 2\ AlBr_3(s)$ c. $Hg(l) + I_2(s) \rightarrow HgI_2(s)$
       f. $Ni(s) + F_2(g) \rightarrow NiF_2(s)$

11. There is only one reactant in a decomposition reaction.

12. a. Hydrogen gas and oxygen gas are the products in the decomposition of water.
    b. Electricity is required to make the decomposition of water proceed.
    c. The process is called electrolysis.

13. **Textbook questions 5.a. and 5.d. of "Practice Problem," p. 94**

    5. a. $8\ MgS(s) \rightarrow 8\ Mg(s) + S_8(s)$ d. $NiCl_2(s) \rightarrow Ni(s) + Cl_2(g)$

14. Gasoline, candle wax, plastics, and methane are examples of hydrocarbons.

15. $x = 1$ and $y = 4$

16. a. Carbon dioxide and water are the products of the complete combustion of a hydrocarbon.
    b. Plenty of oxygen (excess oxygen) is required for the complete combustion of a hydrocarbon.
    c. $CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

17. Figure A3.19 shows that the combustion reaction in an oxy-acetylene torch produces very high temperatures—hot enough to cut steel.

18. **Textbook questions 6.b. and 6.c. of "Practice Problem," p. 95**

    6. b. $2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 4\ CO_2(g) + 6\ H_2O(g)$
       c. $C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$

    Hint: Balance the carbon and hydrogen atoms first. Do the oxygen last since it is by itself.

19. a. Silver is replaced by magnesium when magnesium metal reacts with silver nitrate.

    b. Two silver atoms are replaced by one magnesium atom because the charge on the silver ion is 1+, whereas the charge on the magnesium ion is 2+. It takes two silver ions to make the same charge as one magnesium ion.

20. The charge on the copper ion is 2+, whereas the charge on the aluminium ion is 3+. It takes three copper ions to make the same charge as two aluminium ions.

21. **Textbook question 8 of "Practice Problems," p. 97**

    8. **Word equation:** solid zinc + aqueous silver nitrate → aqueous zinc nitrate + solid silver

       Skeleton equation: $Zn(s) + AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + Ag(s)$

       Balanced equation: $Zn(s) + 2\ AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2\ Ag(s)$

22. a. Bromine and iodine exchange in the reaction. These elements are non-metals.
    b. Six atoms of bromine are required to replace six iodide ions in iron(III) iodide.

23. **Textbook question 7 of "Practice Problems," p. 97**

    7. **Word equation:** chlorine gas + aqueous nickel(III) bromide → aqueous nickel(III) chloride + liquid bromine

       Skeleton equation: $Cl_2(g) + NiBr_3(aq) \rightarrow NiCl_3(aq) + Br_2(l)$

       Balanced equation: $3\ Cl_2(g) + 2\ NiBr_3(aq) \rightarrow 2\ NiCl_3(aq) + 3\ Br_2(l)$

24. **Textbook questions 1.a., 1.e., 2.c., 2.f., 3.b., and 3.d. of "Skill Practice: Decomposition and Single Replacement Reactions," p. 97**

    1. a. magnesium + phosphorus

    2. c. $Mg_3N_2(s) \rightarrow Mg(s) + N_2(g)$

    3. b. $2\ Cu_2O(s) \rightarrow 4\ Cu(s) + O_2(g)$

    e. copper + aluminium iodide

    f. $Pb(s) + AgNO_3(aq) \rightarrow Ag(s) + Pb(NO_3)_2(aq)$

    d. $3\ Br_2(l) + 2\ CrI_3(aq) \rightarrow 2\ CrBr_3(aq) + 3\ I_2(s)$

25. a. **Textbook questions 1 to 8 of "Analyzing and Interpreting," p. 99**

1. Adding water to anhydrous copper(II) sulfate results in a noticeable increase in heat and a colour change from light blue to dark blue.

2. The reaction is exothermic because heat is released.

3. The steel wool burns much faster and much more brightly in pure oxygen than in air.

4. $4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)$

5. There is condensation around the inside of the upper part of the test tube and some steam escaping from the mouth of the test tube.

6. The second product is copper(II) sulfate. This is copper(II) sulfate pentahydrate with the water removed. (Pentahydrate means 5 H_2O.)

7. Both metals grew crystals in the shape of needles. Some of the needles were longer than others. The metals differ in the rate at which the crystals grew. The crystals grew more rapidly with the magnesium than with the copper. No crystals grew with the silver metal.

8. a. $Mg(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Mg(NO_3)_2(aq)$
   b. $Cu(NO_3)_2(aq) + Mg(s) \rightarrow Mg(NO_3)_2(aq) + Cu(s)$
   c. $2\ AgNO_3(aq) + Cu(s) \rightarrow 2\ Ag(s) + Cu(NO_3)_2(aq)$

b. **Textbook question 9 of "Forming Conclusions," p. 99**

9. Your summary should be similar to the following.

   In Part 1, the release of heat and a colour change from light blue to dark blue indicated that a chemical reaction took place.

   In Part 2, the bright glow of the steel wool indicated that a reaction had occurred. The steel wool glowed more brightly in pure oxygen, indicating that the reaction was faster (or more intense).

   In Part 3, water vapour condensed around the inside edge of the test tube and some steam escaped. Also, the colour of the copper(II) sulfate pentahydrate changed from dark blue to light blue.

   In Part 4, the growth of crystals in the solution indicated that a reaction occurred. Both magnesium and copper metal grew crystals with silver nitrate solution, indicating that a reaction occurred. The silver metal and copper(II) nitrate solution did not grow any crystals, indicating that no reaction had occurred.

26. The formation of a precipitate is the evidence that often occurs in a double replacement reaction.

27. No, positive ions can only pair up with negative ions and vice versa.

28. The positive ions always appear first in the chemical formulas of compounds in a double replacement reaction.

29. a. In Example Problem A3.9, lead(II) iodide formed a precipitate because this product is only slightly soluble in water.

    b. Sodium nitrate did not form a precipitate because it is very soluble in water.

    c. Yes, according to Table A2.14, compounds formed with iodide ions (I^-) and lead(II) ions (Pb^{2+}) are only slightly soluble. If a substance is only slightly soluble, it will form a precipitate. Also, according to Table A2.14, compounds formed with nitrate ions (NO_3^-) and all metal ions are very soluble.

30. The purpose of this activity is to observe the formation of precipitates in a number of reactions and to write the balanced chemical equations that represent these reactions.

31. Your completed table should be similar to the following.

	Silver Nitrate	Sodium Hydroxide	Calcium Chloride	Sodium Carbonate
Sodium Iodide	white precipitate		no precipitate	
Iron(III) Chloride		pale yellow precipitate		
Sodium Carbonate			white precipitate	
Silver Nitrate				tan precipitate

32. **Textbook question 1 of "Questions," p. 101**

    1. Sodium iodide and silver nitrate produced a white precipitate.

       Word equation: sodium iodide + silver nitrate → sodium nitrate + silver iodide

       Skeleton equation: $NaI(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgI(s)$

       Balanced equation: $NaI(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgI(s)$

Iron(III) chloride and sodium hydroxide produced a pale yellow precipitate.

Word equation: iron(III) chloride + sodium hydroxide → sodium chloride + iron(III) hydroxide

Skeleton equation: $FeCl_3(aq) + NaOH(aq) \rightarrow NaCl(aq) + Fe(OH)_3(s)$

Balanced equation: $FeCl_3(aq) + 3\ NaOH(aq) \rightarrow 3\ NaCl(aq) + Fe(OH)_3(s)$

Sodium carbonate and calcium chloride produced a white precipitate.

Word equation: sodium carbonate + calcium chloride → sodium chloride + calcium carbonate

Skeleton equation: $Na_2CO_3(aq) + CaCl_2(aq) \rightarrow NaCl(aq) + CaCO_3(s)$

Balanced equation: $Na_2CO_3(aq) + CaCl_2(aq) \rightarrow 2\ NaCl(aq) + CaCO_3(s)$

Silver nitrate and sodium carbonate produced a tan precipitate.

Word equation: silver nitrate + sodium carbonate → sodium nitrate + silver carbonate

Skeleton equation: $AgNO_3(aq) + Na_2CO_3(aq) \rightarrow NaNO_3(aq) + Ag_2CO_3(s)$

Balanced equation: $2\ AgNO_3(aq) + Na_2CO_3(aq) \rightarrow 2\ NaNO_3(aq) + Ag_2CO_3(s)$

33. Your completed table should be similar to the following.

Type of Reaction	Reactants	Products
formation (synthesis)	A + B	AB
decomposition	AB	A + B
hydrocarbon combustion	$C_xH_y + O_2$	$CO_2 + H_2O$
single replacement	A + BC	B + AC
	D + BC	C + BD
double replacement	AB + CD	AD + CB

34. a. It was predicted to be a decomposition reaction because there is only one reactant. According to Table A3.3, a decomposition reaction only has one reactant that is a compound.

b. You can predict that lead will be solid and bromine a liquid because lead is a solid and bromine a liquid at room temperature.

35. a. It was predicted to be a double replacement reaction because two ionic compounds are shown to be reacting.

    b. You know that the products will be nickel(III) sulfite and sodium nitrate because, according to Table A3.3, the compounds in a double replacement reaction exchange ions.

36. **Textbook questions 10.a., 10.c., and 10.d. of "Practice Problem," p. 103**

    10. a. **Reaction type:** hydrocarbon combustion reaction (A hydrocarbon combines with oxygen.)

        Product names: carbon dioxide and water (always the products of complete hydrocarbon combustion)

        Skeleton equation: $C_4H_{10}(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

    c. **Reaction Type:** single replacement reaction (An element combines with an ionic compound.)

        Product names: calcium nitrate and silver metal

        Skeleton equation: $Ca(s) + AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + Ag(s)$

    d. **Reaction Type:** formation (or synthesis) reaction (Two elements combine.)

        Product name: magnesium oxide

        Skeleton equation: $Mg(s) + O_2(g) \rightarrow MgO(s)$

37. The purpose is to classify a variety of chemical reactions and to write the chemical equations that represent them.

38. Observations at each step were as follows:

    step 1: The aluminium foil begins to corrode and bubbles appear. A brown solid appears and sinks to the bottom. This brown solid is copper. The system heats up slightly. **Note:** The brown solid may appear as brown spots on the aluminium foil and the blue-green colour of the copper(II) chloride may disappear.

    step 2: The colbalt(II) chloride turns from pink (pale red) to blue.

    step 3: After a short period of time (10 s to 30 s), you can see crystals of silver metal growing on the copper metal. The solution turns green as copper(II) ions form. (See the photograph of solid silver crystals on copper wire in the Photo Gallery for Module 1 on the Science 10 Multimedia CD.)

    step 4: A white precipitate forms.

39. **Textbook questions 1 to 5 of "Questions," p. 104**

1. **Word equation:** copper(II) chloride + aluminium → aluminium chloride + copper

   Balanced equation: $3\ CuCl_2(aq) + 2\ Al(s) \rightarrow 2\ AlCl_3(aq) + 3\ Cu(s)$

2. The cobalt(II) chloride hexahydrate is a decomposition reaction. The products are anhydrous cobalt(II) chloride and water. Anhydrous cobalt(II) chloride is blue, and you should have seen some water condense on the walls of the test tube. **Note:** The definition of AB → A + B was for simple decompositioin reactions. This decomposition reaction produces two compounds.

3. The reaction of silver nitrate and copper metal is a single replacement reaction. (A compound and an element react.) The solid is silver metal.

4. The reaction involving calcium chloride and sodium carbonate is a double replacement reaction. (Two compounds react to form two new compounds.)

   Word equation: calcium chloride + sodium carbonate → calcium carbonate + sodium chloride

   Balanced equation: $CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2\ NaCl(aq)$

   The white precipitate is calcium carbonate.

5. A formation reaction and a hydrocarbon combustion reaction did not occur in this activity.

40. You should never point the open end of a test tube toward someone while heating a substance in it because the contents may overheat and eject some material out the end of the test tube.

41. **Textbook question 11 of "Practice Problems," p. 105**

11. **Word equation:** lead(IV) nitrate + zinc metal → zinc nitrate + lead

    Skeleton equation: $Pb(NO_3)_4(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + Pb(s)$

    Balanced equation: $Pb(NO_3)_4(aq) + 2\ Zn(s) \rightarrow 2\ Zn(NO_3)_2(aq) + Pb(s)$

42. **Textbook questions 1.b., 1.e., 1.g., 1.j., 2.c., 2.g., 3.b., 3.d, 5, and 6 of "Check and Reflect," p. 106**

1. b. This is a single replacement reaction. (One element in the compound is replaced by another.)

   Balanced equation: $Mg(ClO_4)_2(aq) + 2\ Na(s) \rightarrow 2\ NaClO_4(aq) + Mg(s)$

   Note: $Mg(ClO_4)_2$ and $NaC1O_4$ must be solutions in this reaction.

e. This is a hydrocarbon combustion reaction. (A hydrocarbon reacts with oxygen to form carbon dioxide and water.)

Balanced equation: $2\ C_5H_{10}(l) + 15\ O_2(g) \rightarrow 10\ CO_2(g) + 10\ H_2O(g)$

g. This is a decomposition reaction. (A compound is broken down into its elements.)

Balanced equation: $PbO_2(s) \rightarrow Pb(s) + O_2(g)$

j. This is a double replacement reaction. (Two compounds react to form two new compounds.)

Balanced equation: $3\ NaI(aq) + AlCl_3(aq) \rightarrow\ 3\ NaCl(aq) + AlI_3(s)$

2. c. This is a double replacement reaction. (Two compounds react to form two new compounds.)

Balanced equation: $Sr(OH)_2(aq) + PbBr_2(s) \rightarrow SrBr_2(aq) + Pb(OH)_2(s)$

g. This is a formation reaction. (Two elements react to form a compound.)

Balanced equation: $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$

3. b. This is a decomposition reaction. (A compound is broken up into its elements.) The products are Cu(s) and $Cl_2(g)$.

d. This is a double replacement reaction. (Two compounds react to form two new compounds.) The products are $Ca(NO_3)_2(aq)$ and $PbBr_2(s)$.

5. $2\ HgO(s) \rightarrow 2\ Hg(l) + O_2(g)$

6. $2\ C_6H_6(l) +\ 15\ O_2(g) \rightarrow 12\ CO_2(g) + 6\ H_2O(g)$

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Lesson 4

The Mole

You know that this carton contains a dozen eggs. A dozen is also used for twelve of anything. A dozen oranges is twelve oranges; a dozen pencils is twelve pencils. Chemists need to be able to measure quantities of atoms, ions, molecules, or any other type of particles of substances. Since an atom or molecule is very small, the number of atoms or molecules of a substance in a measured amount needs to be very large.

mole: a quantity used by chemists to measure amounts of elements and compounds

Avogadro's number: the number of particles in a mole of any substance

Chemists have chosen a quantity they call the **mole** to represent a group of atoms or molecules of a substance. In chemistry, a mole is a very large number known as **Avogadro's number**. A mole of any chemical contains Avogadro's number of atoms or molecules.

Avogadro's Number

6.02×10^{23}

or

602 000 000 000 000 000 000 000

Turn to page 107 of the textbook and read the introductory paragraphs of "The Mole." Then continue on by reading "Avogadro's Number and the Mole" on pages 107 and 108.

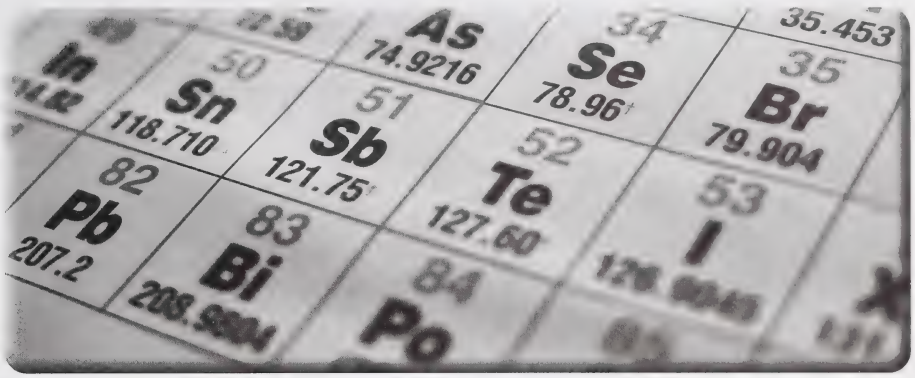
1. Why was carbon-12 chosen to establish the basis of the mole?
2. What is the mass of one mole of carbon-12?
3. What kinds of particles can the mole be used to measure?
4. Why is a mole a convenient measure of most chemicals for chemists?

Check your answers with those on page 162.

molar mass: the mass of one mole of a substance

atomic molar mass: the average mass of one mole of an element's atoms including all isotopes of the element

Chemists have defined **molar mass** and **atomic molar mass** as the mass of one mole of a given substance. Turn to page 108 of the textbook and read the first two paragraphs of "Molar Mass."



5. What is molar mass? What is atomic molar mass?

Check Check your answer with the one on page 162.

The atomic molar mass is given for each element in the periodic table. Turn to page 30 or page 488 of the textbook to reference the periodic table. You can read the atomic molar mass for each element below the element's name.

Example 3.2

What is the atomic molar mass of oxygen?

Solution

The atomic molar mass of oxygen is 16.00 g/mol.

6. State the atomic molar mass of each of the following substances.

   a. hydrogen

   b. potassium

Check Check your answers with those on page 162.

Previously in this section, you wrote chemical formula equations for various reactions. Using the atomic molar mass for each element, you can determine the molar mass for each of the formula units you used in the equations.

Review the rules for significant digits by reading "Calculations with Significant Digits" on pages 467 and 468 of the textbook. The rules for significant digits when adding or subtracting differ from the rules for significant digits when multiplying or dividing. Work through Example 6.1, 6.2, and 6.3, and pay attention to significant digits while doing problems involving molar mass and number of moles.

Now, return to page 108 of the textbook and read the remaining information in "Molar Mass," working through Example Problem A3.12.

7. a. Why is the atomic molar mass of hydrogen multiplied by 4?
   b. How is the molar mass of methane obtained?

8. Answer questions 13, 15, and 16 of "Practice Problems" on page 108 of the textbook.

9. What equation relates the number of moles of a substance to its molar mass?

Check your answers with those on pages 162 and 163.

Often, you need to find the number of moles in a sample of a substance that is less than or more than one mole. In other situations, you may need to find the mass of several moles of a substance. Although the equation $m = n \times M$ can be manipulated to determine n or M, it is usually more understandable to use a method that shows the units involved. This is called the factor-label method.

Read "The Factor-Label Method of Converting Between Quantities" on pages 108 and 109 of the textbook. Work through Example Problems A3.13 and A3.14.

10. In Example Problem A3.13, which factor is used? What unit do you have left after the cancellation step?

11. In Example Problem A3.14, which factor is used? What unit do you have left after the cancellation step?

12. Answer questions 17, 18, and 20 of "Practice Problems" on page 109 of the textbook.

Check your answers with those on pages 163 and 164.

In the next activity you will investigate the relationship between the moles of one substance (iron) that are consumed and the moles of another substance (copper) that are produced during a chemical reaction.

Inquiry Lab

Moles of Copper and Iron

Turn to pages 110 and 111 of the textbook and read the entire activity. Review the list of materials and equipment, and pay particular attention to the steps outlined in the procedure.

If you have access to a supervised laboratory, do **Part A**. If you do not have access to a supervised laboratory, do **Part B**.

Part A

Obtain the required materials and equipment listed on page 110 of the textbook, and complete the procedure. **Pay special attention to the safety precautions mentioned and the hazards noted for each chemical.**

13. Copy and complete a table similar to the following.

Mass of Beaker	Mass of Copper(II) Chloride	Mass of Both Iron Nails		Mass of Beaker and Copper After Reaction
		Before Reaction	After Reaction	

14. Answer the following on pages 110 and 111 of the textbook.

    a. questions 1 to 6 of "Analyzing and Interpreting"
    b. question 7 of "Forming Conclusions"

Check your answers with those on pages 164 and 165.

Part B

Insert the Science 10 Multimedia CD into your computer, and view the segment "Moles of Copper and Iron." Use the information provided in this segment to answer questions 13 and 14 in Part A. **Note:** In question 14.a., you only need to answer questions 1 to 5 of "Analyzing and Interpreting."

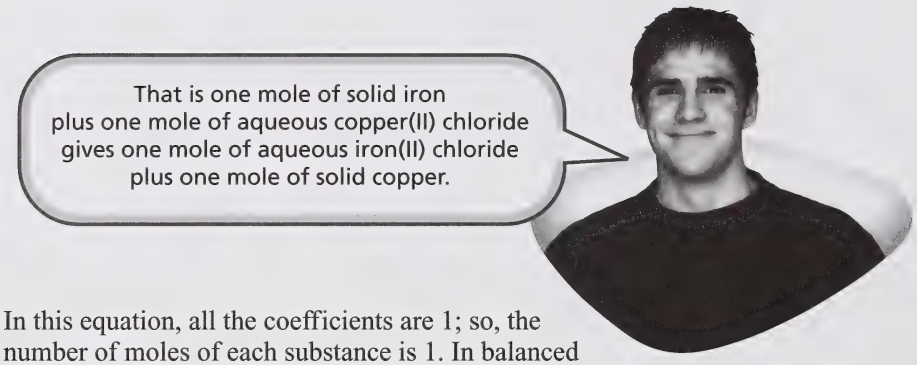
In the previous activity you found that one mole of copper was produced for every mole of iron used in the chemical reaction of iron and copper(II) chloride. The balanced chemical equation shows this 1:1 relationship as well.

$$Fe(s) + CuCl_2(aq) \rightarrow FeCl_2(aq) + Cu(s)$$

Notice that the ratio of formula units of iron to formula units of copper in the balanced equation is 1:1.

When chemists use balanced chemical equations, they read the coefficients as moles of the substance. Using the equation from the previous Inquiry Lab, this would read as follows:

$$1 \text{ mole of } Fe(s) + 1 \text{ mole of } CuCl_2(aq) \rightarrow 1 \text{ mole of } FeCl_2(aq) + 1 \text{ mole of } Cu(s)$$

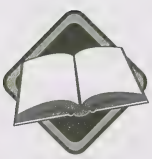

That is one mole of solid iron
plus one mole of aqueous copper(II) chloride
gives one mole of aqueous iron(II) chloride
plus one mole of solid copper.

In this equation, all the coefficients are 1; so, the number of moles of each substance is 1. In balanced equations where the coefficients include numbers other than 1, the number of moles of that substance will be the coefficients of the formula unit of the substance. For example, look at the following equation:

$$3\ NaI(aq) + AlCl_3(aq) \rightarrow 3\ NaCl(aq) + AlI_3(s)$$

In this balanced equation, it reads

$$3 \text{ moles of } NaI(aq) + 1 \text{ mole of } AlCl_3(aq) \rightarrow 3 \text{ moles of } NaCl(aq) + 1 \text{ mole of } AlI_3(s)$$



Turn to page 111 of the textbook and read "The Mole Concept and the Law of Conservation of Mass."

15. How does a balanced equation involving atoms and molecules or moles show that mass is conserved?

Check your answer with the one on page 165.

Read the information given in "*re*SEARCH" on page 111 of the textbook. Visit the Internet site provided to find out more about Amedco Avogadro.

Did You Know?

If you had Avogadro's number of toonies, or 1 mol of toonies, they would cover all of Canada to a height of about 60 km and all of Earth to a height of about 1 km.

Looking Back

You have now completed the concepts of this lesson. You defined the mole in terms of Avogadro's number of particles and described molar mass and atomic molar mass. You then calculated the number of moles of a substance and the mass of a substance given the number of moles of that substance. You also analyzed the mole concept in terms of the law of conservation of mass.

16. Answer questions 1, 4, 5.a., 5.b., 6.a., 6.e., 7.c., 7.e., 8, 9, and 12 of "Check and Reflect" on page 112 of the textbook.

Check your answers with those on pages 165 and 166.

Go to pages 5 to 7 of Assignment Booklet 1C and answer questions 19 to 27.

Glossary

atomic molar mass: the average mass of one mole of an element's atoms including all isotopes of the element

Avogadro's number: the number of particles in a mole of any substance

A mole is 6.02×10^{23} particles.

molar mass: the mass of one mole of a substance

mole: a quantity used by chemists to measure amounts of elements and compounds

Suggested Answers

1. Carbon-12 was chosen as the basis of a mole because it is a solid that is stable and can be obtained in pure form.

2. The mass of one mole of carbon-12 is 12.0 g.

3. The mole can be used to measure any kind of particle—atoms, ions, molecules, and formula units.

4. A mole of most substances will fit into a small beaker, thus making it a convenient quantity for chemists to work with.

5. Molar mass is the mass of one mole of a substance.

   Atomic molar mass is the average mass of one mole of an element's atoms, including all the isotopes of the element.

6. a. The atomic molar mass of hydrogen is 1.01 g/mol.
   b. The atomic molar mass of potassium is 39.10 g/mol.

7. a. The atomic molar mass of hydrogen is multiplied by 4 because there are four atoms of hydrogen in methane.

   b. The molar mass of methane is obtained by adding the molar mass of carbon to the molar mass of the four hydrogen atoms.

8. **Textbook questions 13, 15, and 16 of "Practice Problems," p. 108**

   13. One formula unit of $CH_3OH(l)$ contains one atom of carbon, one atom of oxygen, and four atoms of hydrogen.

   $$\begin{aligned} \text{atomic molar mass of C} &= 12.01\ \text{g/mol} \times 1 = 12.01\ \text{g/mol} \\ \text{atomic molar mass of O} &= 16.00\ \text{g/mol} \times 1 = 16.00\ \text{g/mol} \\ \text{atomic molar mass of H} &= 1.01\ \text{g/mol} \times 4 = 4.04\ \text{g/mol} \\ \text{molar mass of } CH_3OH(l) &= M = 32.05\ \text{g/mol} \end{aligned}$$

   The molar mass of $CH_3OH(l)$ is 32.05 g/mol.

   15. One formula unit of $CO_2(g)$ contains one atom of carbon and two atoms of oxygen.

   $$\begin{aligned} \text{atomic molar mass of C} &= 12.01\ \text{g/mol} \times 1 = 12.01\ \text{g/mol} \\ \text{atomic molar mass of O} &= 16.00\ \text{g/mol} \times 2 = 32.00\ \text{g/mol} \\ \text{molar mass of } CO_2(g) &= M = 44.01\ \text{g/mol} \end{aligned}$$

   The molar mass of $CO_2(g)$ is 44.01 g/mol.

16. One formula unit of $(NH_4)_3PO_4(s)$ contains three atoms of nitrogen, twelve atoms of hydrogen, one atom of phosphorus, and four atoms of oxygen.

atomic molar mass of N	$= 14.01 \text{ g/mol} \times 3 =$	42.03 g/mol
atomic molar mass of H	$= 1.01 \text{ g/mol} \times 12 =$	12.12 g/mol
atomic molar mass of P	$= 30.97 \text{ g/mol} \times 1 =$	30.97 g/mol
atomic molar mass of O	$= 16.00 \text{ g/mol} \times 4 =$	64.00 g/mol
molar mass of $(NH_4)_3PO_4(s)$ =	$M =$	149.12 g/mol

The molar mass of $(NH_4)_3PO_4(s)$ is 149.12 g/mol.

9. The equation that relates the number of moles of a substance to its molar mass is $m = n \times M$, where m is the mass (quantity) of the substance, n is the number of moles of matter, and M is the molar mass of the substance.

10. The factor used is $\frac{1 \text{ mol}}{28.09 \text{ g}}$. The unit remaining is mol, which is what was asked for in the question.

11. The factor used is $\frac{18.02 \text{ g}}{1 \text{ mol}}$. The unit remaining after cancellation is g (grams), which is what was asked for in the question.

12. **Textbook questions 17, 18, and 20 of "Practice Problems," p. 109**

17. **step 1:** Determine the molar mass of NaOH(s).

atomic molar mass of Na	$= 22.99 \text{ g/mol} \times 1 =$	22.99 g/mol
atomic molar mass of O	$= 16.00 \text{ g/mol} \times 1 =$	16.00 g/mol
atomic molar mass of H	$= 1.01 \text{ g/mol} \times 1 =$	1.01 g/mol
molar mass of NaOH(s) =	$M =$	40.00 g/mol

step 2: The factor is $\frac{40.00 \text{ g}}{1 \text{ mol}}$.

$$\text{Therefore, } m_{NaOH} = 5.0 \cancel{\text{mol}} \times \frac{40.00 \text{ g}}{1 \cancel{\text{mol}}}$$

$$= 2.0 \times 10^2 \text{ g} \quad \leftarrow \text{2 significant digits}$$

The mass of 5.0 mol of NaOH(s) is 2.0×10^2 g. (**Note:** The final answer has only 2 significant digits because 5.0 mol has only 2 significant digits. If you wrote the answer as 200 g, it would show 3 significant digits. This is incorrect.)

18. **step 1:** Determine the molar mass of $C_6H_{12}O_6(s)$.

$$\begin{aligned}
\text{atomic molar mass of C} &= 12.01\ \text{g/mol} \times 6 = 72.06\ \text{g/mol} \\
\text{atomic molar mass of H} &= 1.01\ \text{g/mol} \times 12 = 12.12\ \text{g/mol} \\
\text{atomic molar mass of O} &= 16.00\ \text{g/mol} \times 6 = 96.00\ \text{g/mol} \\
\text{molar mass of } C_6H_{12}O_6(s) &= M = 180.18\ \text{g/mol}
\end{aligned}$$

step 2: The factor is $\frac{1\ \text{mol}}{180.18\ \text{g}}$.

$$\text{Therefore, } n_{C_6H_{12}O_6} = 360\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{180.18\ \cancel{\text{g}}}$$
$$= 2.00\ \text{mol} \quad \leftarrow \text{3 significant digits}$$

There are 2.00 mol in a 360-g sample of $C_6H_{12}O_6(s)$.

20. **step 1:** Determine the molar mass of $Mg(NO_3)_2(aq)$.

$$\begin{aligned}
\text{atomic molar mass of Mg} &= 24.31\ \text{g/mol} \times 1 = 24.31\ \text{g/mol} \\
\text{atomic molar mass of N} &= 14.01\ \text{g/mol} \times 2 = 28.02\ \text{g/mol} \\
\text{atomic molar mass of O} &= 16.00\ \text{g/mol} \times 6 = 96.00\ \text{g/mol} \\
\text{molar mass of } Mg(NO_3)_2(aq) &= M = 148.33\ \text{g/mol}
\end{aligned}$$

step 2: The factor is $\frac{1\ \text{mol}}{148.33\ \text{g}}$.

$$\text{Therefore, } n_{Mg(NO_3)_2} = 20.0\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{148.33\ \cancel{\text{g}}}$$
$$= 0.135\ \text{mol} \quad \leftarrow \text{3 significant digits}$$

There are 0.135 mol of $Mg(NO_3)_2(aq)$ in a 20.0-g sample.

13. Your completed table should be similar to the following.

Mass of Beaker	Mass of Copper(II) Chloride	Mass of Both Iron Nails		Mass of Beaker and Copper After Reaction
		Before Reaction	After Reaction	
58.40 g	5.34 g	9.31 g	8.61 g	59.62 g

14. a. **Textbook questions 1 to 6 of "Analyzing and Interpreting," pp. 110 and 111**

1. mass of iron used = mass of iron before reaction − mass of iron after reaction
   = 9.31 g − 8.61 g
   = 0.70 g

2. mass of copper produced = mass of beaker and dried copper − mass of beaker
   = 59.62 g − 58.40 g
   = 1.22 g

3. $n_{Fe} = 0.70 \text{ g} \times \frac{1 \text{ mol}}{55.85 \text{ g}}$

$= 0.013 \text{ mol}$

4. $n_{Cu} = 1.22 \text{ g} \times \frac{1 \text{ mol}}{63.55 \text{ g}}$

$= 0.0192 \text{ mol}$

5. ratio of moles of iron to moles of copper $= \frac{0.013 \text{ mol}}{0.0192 \text{ mol}}$

$= 0.68$ ← rounded to the nearest whole number

The ideal ratio is 1:1. The calculated ratio of 0.68 (to 2 significant digits) shows that there is some error involved. You may have obtained a result closer to the ideal ratio.

6. Students should have a ratio close to 1:1.

b. **Textbook question 7 of "Forming Conclusions," p. 111**

7. In this reaction, one mole of copper is produced for every mole of iron used. The balanced equation for this reaction is as follows:

$$Fe(s) + CuCl_2(aq) \rightarrow Cu(s) + FeCl_2(aq)$$

15. The balanced equations involving atoms and molecules or moles has the same number of atoms, molecules, or moles of each substance on both sides of the equation.

16. **Textbook questions 1, 4, 5.a., 5.b., 6.a., 6.e., 7.c., 7.e., 8, 9, and 12 of "Check and Reflect," p. 112**

1. There are 6.02×10^{23} particles in a mole of any substance.

4. The number of particles in one mole of a substance is called Avogadro's number. The symbol is N_A.

5. a. $N_{Au} = 1.0 \text{ mol} \times 6.02 \times 10^{23} \text{ atoms/mol}$
$= 6.0 \times 10^{23} \text{ atoms}$

b. $N_{He} = 2.5 \text{ mol} \times 6.02 \times 10^{23} \text{ atoms/mol}$
$= 1.5 \times 10^{24} \text{ atoms}$

Note: Significant digits affect the answer. Refer to the rules for significant digits and scientific notation on pages 467 to 469 of the textbook.

6. a. molar mass of Na = 22.99 g/mol

Therefore, $n_{Na} = 28 \text{ g} \times \frac{1 \text{ mol}}{22.99 \text{ g}}$

$= 1.2 \text{ mol}$

e. molar mass of $N_2(g) = 14.01 \text{ g/mol} \times 2$

$= 28.02 \text{ g/mol}$

Therefore, $n_{N_2} = 26.0 \text{ g} \times \frac{1 \text{ mol}}{28.02 \text{ g}}$

$= 0.928 \text{ mol}$

7. c. Determine the molar mass of $H_2O(l)$.

atomic molar mass of H = 1.01 g/mol × 2 = 2.02 g/mol
atomic molar mass of O = 16.00 g/mol × 1 = 16.00 g/mol
molar mass of $H_2O(l)$ = M = 18.02 g/mol

Therefore, $m_{H_2O} = 5.0 \text{ mol} \times \frac{18.02 \text{ g}}{1 \text{ mol}}$

$= 90 \text{ g}$

e. Determine the molar mass of $Al_2S_3(s)$.

atomic molar mass of Al = 26.98 g/mol × 2 = 53.96 g/mol
atomic molar mass of S = 32.07 g/mol × 3 = 96.21 g/mol
molar mass of $Al_2S_3(s)$ = M = 150.17 g/mol

Therefore, $m_{Al_2S_3} = 0.001\,27 \text{ mol} \times \frac{150.14 \text{ g}}{1 \text{ mol}}$

$= 0.191 \text{ g}$

8. a. 3 mol + **1 mol** → 2 mol

b. **6 g** + 28 g → 34 g

c. 1.83×10^{24} molecules + 6.02×10^{23} molecules → **1.20×10^{24} molecules**

9. $n_{Cu} = 3.01 \times 10^{23} \text{ Cu(s) atoms} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ atoms}}$

$= 0.500 \text{ mol}$

molar mass of Cu(s) = 63.55 g/mol

Therefore, $m_{Cu} = 0.500 \text{ mol} \times \frac{63.55 \text{ g}}{1 \text{ mol}}$

$= 31.8 \text{ g}$

The mass of 3.01×10^{23} atoms of copper is 31.8 g.

12. A mole of oxygen gas contains twice as many atoms as a mole of iron because each oxygen molecule contains two oxygen atoms.

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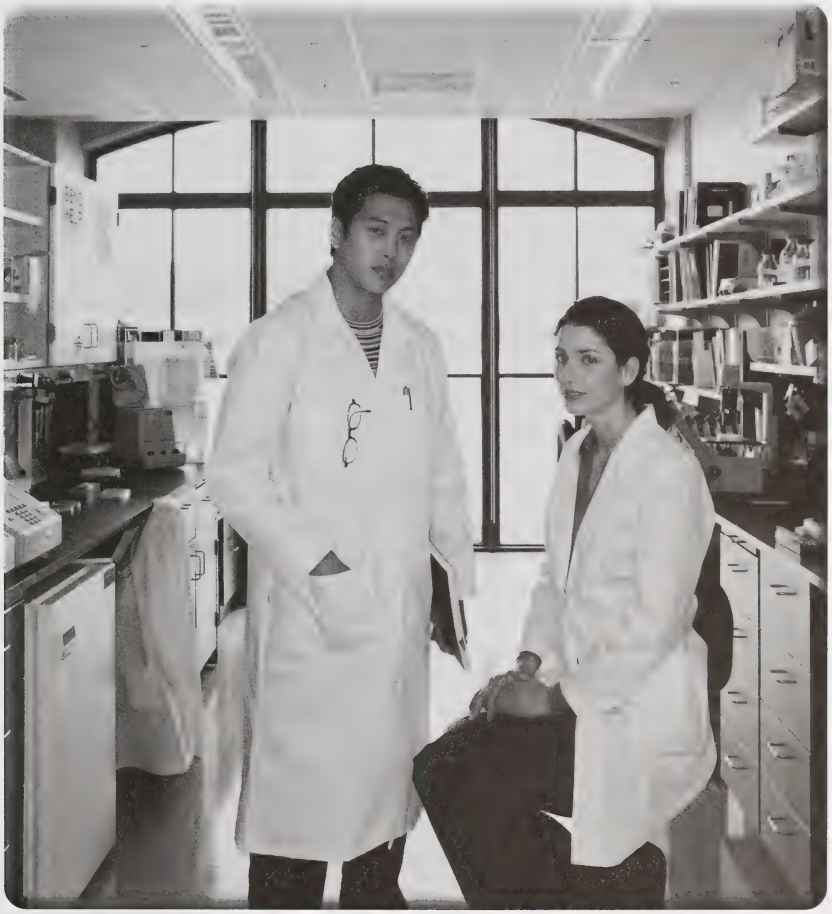
Conclusion

In this section you identified a number of chemical reactions and used the law of conservation of mass to help you to write balanced chemical equations. You described five common types of chemical reactions and wrote balanced chemical equations for each. You then defined the concept of a mole as a quantity used by chemists to measure amounts of chemicals in reactions.

You are now able to write word and balanced chemical equations. You now know the importance of chemical reactions in industry and how an understanding of chemical reactions has allowed for a variety of materials. With all this new knowledge of chemistry, do you think you would be interested in a career as a research chemist?



Module Summary

In this module you studied transformations of matter in chemical changes and how energy is involved.

In Section 1 you described the WHMIS system of hazardous material identification and identified several safe practices regarding the handling of chemicals. You also identified a number of properties of matter and developed some ideas about the structure of matter.

In Section 2 you analyzed the periodic table and described the structure of the atom in terms of a central nucleus and surrounding energy levels. You then identified methods of naming as well as some properties of ionic and molecular compounds. You also identified methods of naming and some properties of acids and bases. Finally, you examined some issues related to chemical use in society and the health concerns that may result from society's dependence on chemicals.

In Section 3 you discovered a number of important examples of chemical change and listed some characteristics of chemical reactions. You observed how to write chemical equations as word equations and formula equations. You then developed a basic competency in balancing chemical equations. You analyzed five common types of chemical reactions and wrote word, skeleton, and balanced equations for these five types of chemical reactions. You summed up this section by describing the concept of the mole and developing a competency in converting between the number of moles and the mass of a substance.

Knowledge of the atomic structure of matter and of some of the chemical properties of matter will help you understand how some of the chemical technologies have advanced.

In Alberta, the oil industry and numerous related chemical production sites provide the raw products for many of the materials and products people take for granted. The chemical industry has also provided career opportunities for a significant portion of the population.

Module Review

Turn to page 116 of the textbook and review the key concepts and some of the things you were presented within this module. Then answer the following "Unit Review" questions on pages 117 to 120. If necessary, go back and read over the parts of this module as you answer the questions.

1. question 1 of "Vocabulary" on page 117

2. questions 5, 15, 16.a., 16.d., 21, 22.a., 22.f., 23.a., and 23.c. of "Knowledge" on pages 117 and 118

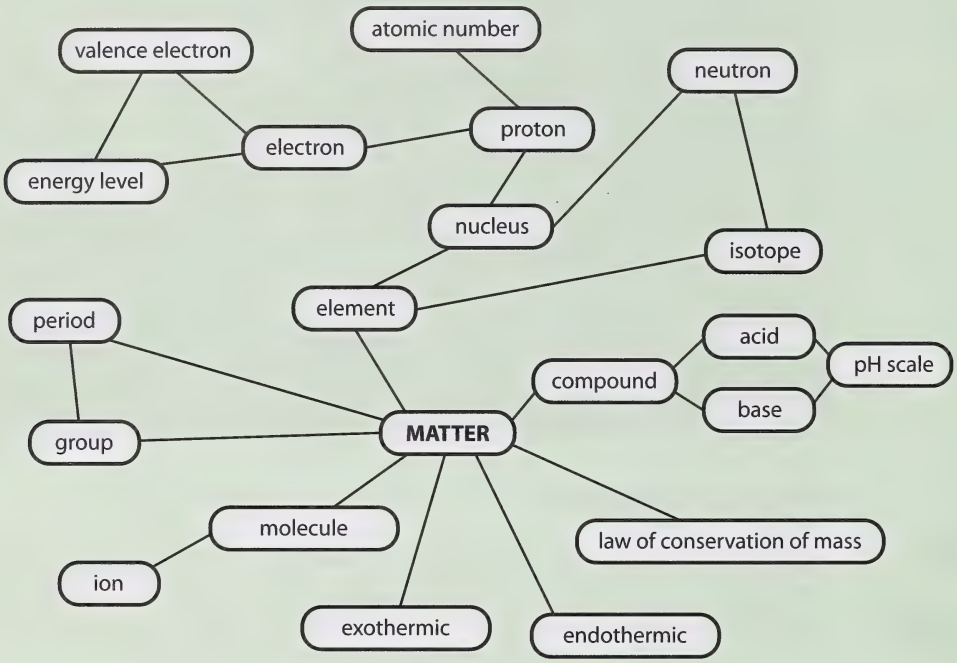
3. questions 25.c., 28, 29, 33.d., 33.h., 34.d., 34.j., 35.c., 36.c., 44, 47.d., 49.c., and 50.c. of "Applications" on pages 118 to 120

Check your answers with those on pages 170 to 172.

Suggested Answers

1. **Textbook question 1 of "Vocabulary," p. 117**

   The following is a sample answer. Your concept map may have words in different positions but the links from one word to the next should be the same.



2. **Textbook questions 5, 15, 16.a., 16.d., 21, 22.a., 22.f., 23.a., and 23.c. of "Knowledge," pp. 117 and 118**

   5. The Inuit found chunks of copper already in metallic form. They used the metallic copper to make arrows, knives, harpoons, rivets, staples, and handles for pots.

15.

Atom or Ion	Overall Charge	Protons	Electrons	Symbol
sulfur atom	0	16	16	S
sulfide ion	2–	16	18	S^{2-}
lithium ion	1+	3	2	Li^{+}
oxide ion	2–	8	10	O^{2-}
chloride ion	1–	17	18	Cl^{-}
iron(II) ion	2+	26	24	Fe^{2+}
nitride ion	3–	7	10	N^{3-}

16. a. very soluble
d. slightly soluble

21. A chemical reaction is a process that occurs when substances react to form different substances. A chemical equation is a representation of a chemical reaction using chemical formulas and symbols.

22. a. $Cl_2(g) + 2\ KBr(aq) \rightarrow 2\ KCl(aq) + Br_2(l)$
f. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

23. a. hydrocarbon combustion reaction
c. single replacement reaction

3. **Textbook questions 25.c., 28, 29, 33.d., 33.h., 34.d., 34.j., 35.c., 36.c., 44, 47.d., 49.c., and 50.c. of "Applications," pp. 118 to 120**

25. c. The safety hazard symbols you would expect to see are flammable and explosive. Because spray paint cans are a consumer item, you would not expect to see any WHMIS symbols on them.

28. a.

b.

29. a. When metals become ions, they **lose** electrons in order to become **positively** charged.
    b. When non-metals become ions, they **gain** electrons in order to become **anions**.

33. d. $Ag_2CO_3(s)$ h. $Co(OH)_3(s)$

34. d. ammonium sulfide j. lead(IV) permanganate

35. c. ClF(g)

36. c. nitrogen trichloride

44. The reaction is exothermic. The beaker becomes warmer because the reaction is releasing heat.

47. d. The reaction is a double replacement reaction.

$$3\ Li_2SO_3(aq) + 2\ Au(NO_3)_3(aq) \rightarrow 6\ LiNO_3(aq) + Au_2(SO_3)_3(s)$$

49. c. Determine the molar mass of $H_2O(l)$.

atomic molar mass of H = 1.01 g/mol × 2 = 2.02 g/mol
atomic molar mass of O = 16.00 g/mol × 1 = 16.00 g/mol
molar mass of $H_2O(l)$ = M = 18.02 g/mol

Therefore, $n_{H_2O} = 72.08\ \cancel{g} \times \frac{1\ mol}{18.02\ \cancel{g}}$

$= 4.000\ mol$ ← 4 significant digits

There are 4.000 mol in a 72.08-g sample of $H_2O(l)$.

50. c. molar mass of $O_2(g) = 16.00\ g/mol \times 2$

$= 32.00\ g/mol$

Therefore, $m_{O_2} = 10\ \cancel{mol} \times \frac{32.00\ g}{1\ \cancel{mol}}$

$= 3.2 \times 10^2\ g$ ← 2 significant digits

The mass of 10 mol of $O_2(g)$ is 3.2×10^2 g.

MATERIAL SAFETY DATA SHEET

SECTION 1—PRODUCT IDENTIFICATION AND USE

PRODUCT IDENTIFIER ▶	Calcium Chloride (Anhydrous) $CaCl_2 \cdot 2H_2O$
PRODUCT IDENTIFICATION NUMBER (PIN)	10043-52-4
PRODUCT USE ▶	

MANUFACTURER'S NAME		SUPPLIER'S NAME	
	Big Chemical Co.		ABC Chemical Inc.
STREET ADDRESS	12345 Somewhere Street	STREET ADDRESS	
CITY	PROVINCE	CITY	PROVINCE
POSTAL CODE	EMERGENCY TELEPHONE NO.	POSTAL CODE	EMERGENCY TELEPHONE NO.

SECTION 2—HAZARDOUS INGREDIENTS

HAZARDOUS INGREDIENTS	%	CAS NUMBER	LD_{50} OF INGREDIENT (SPECIFY SPECIES AND ROUTE)	LC_{50} OF INGREDIENT (SPECIFY SPECIES)
Calcium Chloride	99-100	10043-52-4	oral rat	1000 mg/kg

SECTION 3—PHYSICAL DATA

PHYSICAL STATE	ODOUR AND APPEARANCE			ODOUR THRESHOLD (ppm)
Solid	White granular			N/AV
VAPOUR PRESSURE (mm Hg) N/AP	VAPOUR DENSITY (AIR=1) N/AP	EVAPORATION RATE N/AP	BOILING POINT (°C) >1600°C	FREEZING POINT (°C) 260°C
pH Neutral	SPECIFIC GRAVITY 2.15	COEFF. WATER/OIL DIST. N/AV		

SECTION 4—FIRE AND EXPLOSION DATA

FLAMMABILITY YES ☐ NO ☒ IF YES, UNDER WHICH CONDITIONS? ▶

MEANS OF EXTINCTION

Extingishing media appropriate to surrounding conditions

FLASHPOINT (°C) AND METHOD	UPPER FLAMMABLE LIMIT (% BY VOLUME)	LOWER FLAMMABLE LIMIT (% BY VOLUME)
N/AP	N/AP	N/AP

AUTOIGNITION TEMPERATURE (°C)	HAZARDOUS COMBUSTION PRODUCTS
N/AV	Chlorides

EXPLOSION DATA ▶	SENSITIVITY TO IMPACT	SENSITIVITY TO STATIC DISCHARGE
	Not an explosion hazard	Not a fire hazard

SECTION 5—REACTIVITY DATA

CHEMICAL STABILITY YES ☐ NO ☒ IF NO, UNDER WHICH CONDITIONS? ▶ Will pick up moisture from the air if left in open container

INCOMPATIBILITY WITH OTHER SUBSTANCES YES ☒ NO ☐ IF SO, WHICH ONES? ▶ Water, methyl vinyl ether, zinc, bromine trifluoride, mixtures of lime and boric acid, 2-furan percarboxylic acid

REACTIVITY, AND UNDER WHAT CONDITIONS: Reacts with water to produce considerable amounts of heat

HAZARDOUS DECOMPOSITION PRODUCTS: Emits toxic chlorine fumes when heated to decomposition. May form hydrogen chloride

PRODUCT IDENTIFIER ▶ Calcium Chloride (Anhydrous) $CaCl_2 \cdot 2H_2O$

SECTION 6—TOXICOLOGICAL PROPERTIES

ROUTE OF ENTRY

SKIN CONTACT ☒ SKIN ABSORPTION ☐ EYE CONTACT ☒ INHALATION ☒ INGESTION ☒

EFFECTS OF ACUTE EXPOSURE TO PRODUCT
2-Moderate irritation to skin, eyes, respiratory tract

EFFECTS OF CHRONIC EXPOSURE TO PRODUCT Excessive contact may produce "soda ulcers" on hands and perforation of nasal system. Chronic toxicity not fully investigated.

EXPOSURE LIMITS	IRRITANCY OF PRODUCT	SENSITIZATION TO PRODUCT	CARCINOGENICITY
Oral rat LD_{50}: 1000 mg/kg	Skin, eyes, mucus membrane	No	No
TERATOGENICITY	**REPRODUCTIVE TOXICITY**	**MUTAGENICITY**	**SYNERGISTIC PRODUCTS**
No	No	Investigating	No

SECTION 7—PREVENTIVE MEASURES

PERSONAL PROTECTIVE EQUIPMENT: Body-covering clothing

GLOVES (SPECIFY)	RESPIRATOR (SPECIFY)	EYE (SPECIFY)
Protective gloves	Half-face dust respirator where dust or mist apparent	Safety goggles
FOOTWEAR (SPECIFY)	**CLOTHING (SPECIFY)**	**OTHER (SPECIFY)**
Rubber boots	body-covering clothing	Maintain good housekeeping in work/storage area

ENGINEERING CONTROLS (SPECIFY, EG. VENTILATION, ENCLOSED PROCESS)

Local or general exhaust

LEAK AND SPILL PROCEDURE

Ventilate area, sweep up, and containerize for reclaimation or disposal

WASTE DISPOSAL

According to all applicable regulations. Avoid run-off.

HANDLING PROCEDURES AND EQUIPMENT

Avoid creating product dust. Avoid contact with product.

STORAGE REQUIREMENTS

Store in a cool, dry place, away from flame and sparks. Store in a well-ventilated area. Store in a tightly closed container.

SPECIAL SHIPPING INFORMATION

SECTION 8—FIRST AID MEASURES

SPECIFIC MEASURES

Eyes: Wash with copious quantities of water for at least 15 minutes. Seek immediate medical attention.

Skin: Remove contaminated clothing. Wash skin with soap and water. If irritation occurs or persists, seek medical attention. Wash clothing.

Inhalation: Remove patient to fresh air. Administer approved oxygen supply if breathing is difficult. Administer artifical respiration if breathing has ceased. Call a physician.

Ingestion: Give large quantities of water to dilute. Induce vomiting. Seek immediate medical attention. Never give anything by mouth to an unconscious person.

SECTION 9—PREPARATION DATE OF MSDS

PREPARED BY (GROUP, DEPARTMENT, ETC.)	PHONE NUMBER	DATE

Template for Inquiry Lab: Ionic or Molecular?

Title: _____

Hypothesis: _____

_____

Materials and Apparatus: _____

_____

_____

_____

_____

_____

Procedure: _____

_____

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Observations:

Substance	Appearance	Hardness	Odour	Solubility in Water	Conductivity	Melting Point
sodium chloride						
sugar						
potassium iodide						
bees wax						
Epsom salts						
naphthalene flakes						

Conclusion:

SCIENCE 10

Energy and Matter in Chemical Change

Energy Flow in Technological Systems

Cycling Matter in Living Systems

Energy Flow in Global Systems

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